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PROGRESSIVE EXERCISES

IN

PRACTICAL CHEMISTRY.

BY

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PROFESSOR OF CHEMISTRY IN THE WOMAN'S MEDICAL COLLEGE OF PENNSYLVANIA, IN THE PENNSYLVANIA COLLEGE OF DENTAL SURGERY AND IN THE WAGNER FREE INSTITUTE OF SCIENCE; PATHOLOGICAL CHEMIST TO THE JEFFERSON MEDICAL COLLEGE HOSPITAL.

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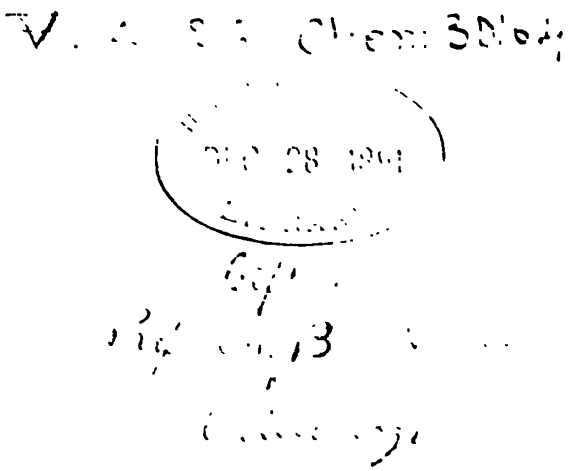

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PREFACE.

The following pages include the methods of laboratory instruction in inorganic chemistry, given during several years in the Woman's Medical College and in the Pennsylvania College of Dental Surgery. Originally carried out by oral demonstrations, the system was subsequently reduced to manuscript, and after some experience, it was decided to put it in book form. The usual method of laboratory instruction, consisting of practice in elementary qualitative analysis, does not seem to us to be adapted to the teaching of chemistry as contemplated in the curriculum of American medical colleges. In our own teaching, we have largely supplanted the lecture hours by laboratory work, and have felt the necessity of progressive exercises in chemical principles, illustrated by experiments safe and easy to perform. We have, therefore, given much attention to details as to quantity of materials to be used and arrangements of apparatus. Some of the experiments and forms of apparatus are new, and have been devised especi-

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ally with a view to economy, a point of much importance when large classes are being taught. Formulæ, and discussions of purely theoretical questions have been omitted, the book being adapted for use in conjunction with any manual of elementary chemical principles, or to be supplemented by lectures.

We are indebted to Dr. Eleanor Moore Hiestand, Demonstrator of Chemistry at the Woman's Medical College, for much assistance in the preparation of the book.

H. L.
W. B.

715 Walnut street.
September, 1890.

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CONTENTS.

APPARATUS AND MANIPULATIONS.

Bunsen Burner—Spirit Lamp—Test Tubes and Accessories—Beakers—Boiling—Mortar—Balance—Tongs—Pouring from Bottles, etc.—Pipette and Burette—Filtration—Wash Bottle—Boring Corks—Bending, Drawing and Cutting Glass Tubes—Evaporation—Distillation, ... 9–22

GENERAL CHEMICAL PRINCIPLES.

Solution—Physical and Chemical Change—Conditions Affecting Chemical Action—Decomposition, Elements and Compounds, ... 23–37

EXERCISES IN CHEMISTRY OF IMPORTANT ELEMENTS.

Hydrogen — Potassium — Sodium — Silver — Chlorine — Bromine — Iodine — Fluorine — Oxygen — Sulphur — Calcium — Barium — Strontium — Lead — Zinc — Magnesium — Copper — Mercury — Aluminum — Iron — Manganese — Chromium — Nickel — Cobalt — Nitrogen — Phosphorus — Arsenic — Antimony — Bismuth — Boron — Gold — Carbon — Silicon — Tin, ... 38–97

ELECTRICAL DECOMPOSITION, ... 98–100

ELEMENTS, ... 101

METRIC WEIGHTS AND MEASURES, ... 102

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ERRATUM.

Page 23, Exp. 2, for "*copper*" read "copper sulphate."

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PROGRESSIVE EXERCISES IN PRACTICAL CHEMISTRY.

APPARATUS AND MANIPULATIONS.

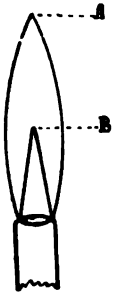
The BUNSEN BURNER is the usual means of employing gas for heating purposes. It yields a non-luminous, smokeless flame. Fig. 1 is the common pattern. The openings at the lower part of the tube permit air to enter, and can be controlled by means of the revolving collar. It may rest on a mat of thin asbestos board about six inches square, which will prevent injury to the table, and be a convenient resting place for hot objects.

FIG. 1.

The principal difficulties in the management of the burner are the "running back" of the flame so as to burn at the jet below, and the stopping of the jet by dirt. "Running back" is caused by low pressure of gas, or by too free air supply. When it occurs the burner will become very hot,

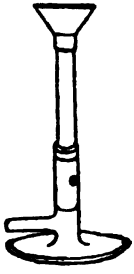
emit a disagreeable odor, and may cause headache. The student should learn to recognize this condition promptly, and should therefore make the following experiment: Turn the gas on full, light the burner, set the collar so that the non-luminous flame is produced, and slowly turn down the gas. A point will generally be reached at which the flame will pass down the tube, often with a slight sound, and can be seen burning below at the jet.

Fig. 2.



If the gas be now turned on full, the odor at once becomes evident. Apply a light at the top of the tube. A thin, feebly luminous flame is produced. When the burner presents this appearance, it should at once be turned out completely, then turned on fully and relighted. To prevent this "running back" the revolving collar should be turned so as to partially close the air openings.

Fig. 3.



In making these adjustments, care should be taken not to burn the fingers; the burner should be handled by means of a piece of cloth or paper.

The Bunsen flame, Fig. 2, consists of an inner and an outer cone, different parts of which possess different chemical and thermal qualities. The hottest part is at the point B, above the top of the inner cone.

The Winged Top, Fig. 3, is used when a spreading flame is required, as in bending glass tubes.

The SPIRIT LAMP, Fig. 4, is of service when moderate heating is required or gas not available. The wick should be kept clean. When not in use the lamp should be covered by the cap, to prevent loss of alcohol.

FIG. 4.

FIG. 5.

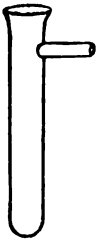
TEST-TUBES, Fig. 5, are small, thin glass tubes, closed at one end, and slightly flared at the mouth. It is convenient to have them about four inches long and in two widths, approximately one-half and three-fourths inch. They are cleaned by the use of the cylindrical brush, Fig. 6.

A very good cleaner may be made by wrapping several layers of muslin around a stick and tying securely with twine. The muslin should project nearly an inch beyond the end of the stick, so as to form a pad.

FIG. 6.

A side-neck tube, Fig. 7, is used in the preparation of gases.

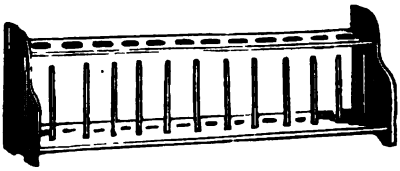
FIG. 7.



Test-tubes are kept in the wooden racks. When in use the tubes stand upright, but clean tubes are places upside down on the pegs. The rack should be washed from time to time. The usual form is shown in Fig. 8.

BEAKERS, Fig. 9, are thin glass vessels with flat bottoms. Two of these will suffice, holding respectively 100 and 150 cc., with a couple of watch-glasses to serve as covers.

FIG. 8.



BOILING is generally performed in test-tubes. Care must be taken to hold the tube in such a position that the flame will not touch the glass above the surface of the liquid, otherwise it may be overheated at this point and fractured. The tube should be held by means of the clamp, Fig. 10, or by the projecting ends of a piece of paper wrapped around it. When the boiling is to be prolonged, the tube can be more conveniently supported in the clamp-stand, Figs. 11 and 12. In order that the coil shall grasp the rod firmly, it must be bent on itself slightly by forcing apart the two middle strands at one side, by a wedge-shaped instrument. This remark applies also to the helix and burette holder described further on.

FIG. 9.

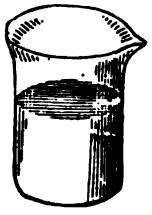
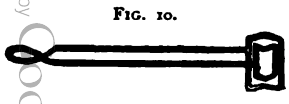


FIG. 10.



To prevent projection of the hot liquid from the tube, the lower end may be rested on the burner, as shown in Fig. 13. In this way the boiling proceeds regularly and quietly.

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Care should be taken that the hands, note-books, or other articles which might be injured by the hot liquid, are not placed so that it can fall upon them, should it boil over or the tube break.

FIG. 11.

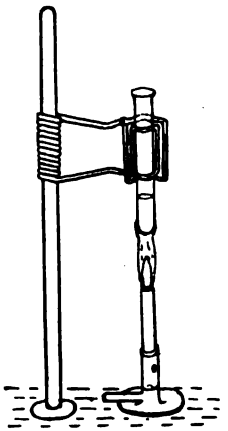
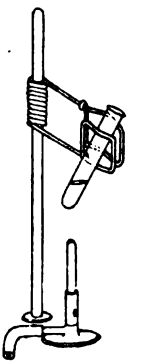


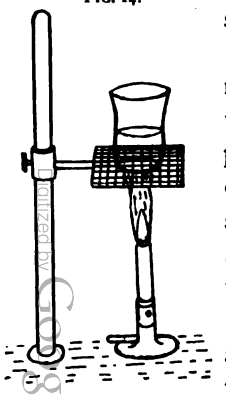
FIG. 12.



Larger volumes of liquid are boiled in beakers, resting upon wire gauze, which distributes the heat and renders breakage less apt to occur, Fig. 14. A plate of iron or sheet of asbestos may be substituted for the gauze.

FIG. 13.

FIG. 14.



MORTAR.—Fig. 15 represents a Wedgewood mortar and pestle used for reducing substances to powder. A small size, three inches across the top, will suffice. The powder may be scraped out of

FIG. 15.

the mortar by means of a steel spatula, Fig. 16. A good spatula may be made by cutting out a strip of thin sheet iron, about ¾ of an inch broad and 3 inches long, and rounding off one end.

FIG. 16

BALANCE.—The ordinary hand-balance, Fig. 17, will serve very well. It should be suspended from the ring-stand, and at such a height that the pans are about half an inch from the table. In order to prevent corrosion of the pans, it is advisable to place on each a piece of white glazed paper.

FIG. 17.

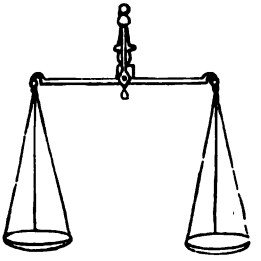
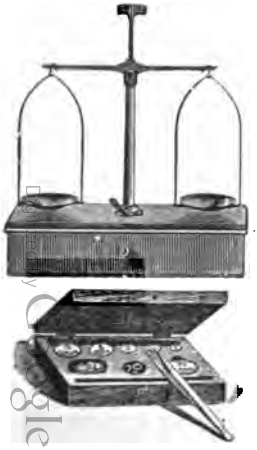


FIG. 18.



Only grain weights are supplied with this balance. For the conversion of grains into grams or *vice versâ*, see Appendix. A more convenient and exact form of balance is shown in Fig. 18.

Fine, dry sea sand may be used for counterpoising vessels. Weights should always be handled by means of the forceps.

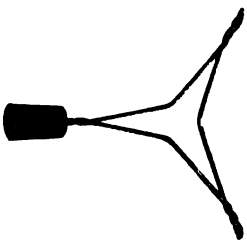
Measurement of Liquids.—Liquid measurement is now almost entirely

made according to the metric (decimal) system, the unit of which is the *cubic centimeter*, usually written cc., which is equal to one gram (15.43 grains) of water at 60 degrees Fahrenheit (15 degrees centigrade).

FIG. 19.

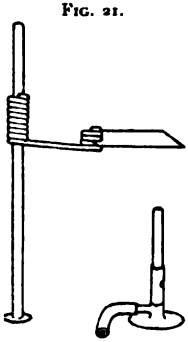
For approximate measurement, a test tube on foot scratched at points corresponding to each five cc. will suffice.

FIG. 20.



TONGS, ETC.—For observing the action of heat on substances, platinum foil may be employed, being held in the iron tongs, Fig. 19. The iron wire triangle, one end of which is thrust through a cork for convenience in handling, Fig. 20, may also be used. Platinum foil should be kept clean by use of hydrochloric acid, and burnishing with fine white sea sand. Lead, tin, and other fusible metals, should not be heated on it. Pieces of thin sheet iron may often be substituted for platinum foil, but cannot usually be employed a second time. They may be held in the helix, Fig. 21.

FIG. 21.



POURING FROM BOTTLES. — In taking solids, such as dry powders, from bottles,

the latter should be held in a nearly horizontal position, and gently rotated until some of the powder falls out. When only a very small quantity is desired, it will be best to pour some out on a clean piece of paper, and make the selection from this. In pouring a liquid, the bottle should be held with the label uppermost, so that none of the liquid may run down and destroy it.

Fig. 22.

Fig. 23.

Mushroom stoppers, Fig. 22, may be placed upside down on the table, but the flat stoppers, Fig. 23, should be taken between the fingers of the hand which is to hold the bottle, as shown in Figs. 24 and 25. The greatest care should be taken to avoid the transfer of stoppers from one bottle to another. Before pouring, the dust which has settled on the lip of the bottle should be wiped off. In the case of strong acids, it is safer not to handle the stopper as shown in the cut, but to grasp it upright with the hand holding the test-tube, otherwise the acid may run down and burn the fingers.

Fig. 24.

Fig. 25.

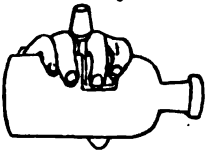
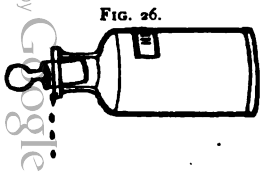


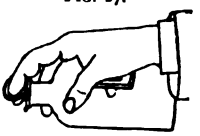
Fig. 26.



Liquids may be made to run out in drops, by first pulling out the stopper a

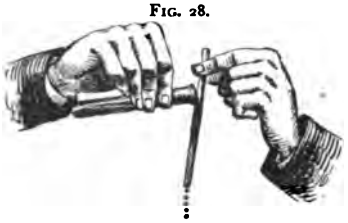
short distance, as shown in Fig. 26; if care be taken the stopper will not fall out. It may be well, until expertness is acquired, to hold the stopper in position with the index and middle fingers, the bottle itself being held by the remaining fingers, as shown in Fig. 27.

FIG. 27.



POURING FROM TEST-TUBES AND BEAKERS.—In pouring a liquid from a test-tube or beaker, a glass rod should be held vertically close to the lip of the vessel, Fig. 28. In this way the liquid is made to follow the rod and does not run down the outside of the vessel.

FIG. 28.

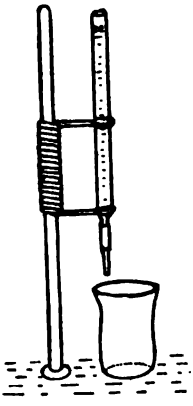


Pipettes, Burettes.—Liquids may be transferred from one vessel to another by means of the pipette, Fig. 29A. It is filled by suction and the liquid retained by pressing the finger on the top of the tube, the flow of liquid being easily controlled by the finger. A *burette* is a pipette provided with a stopcock to control the flow. A cheap and efficient form of stopcock may be made by connecting a glass jet to the end of the pipette by means of a short piece of pure rubber tubing, in which is inserted a glass rod about ⅓ of an inch long and fitting

FIG. 29A.

the tube sufficiently closely to prevent flow of liquid. By pinching the rubber tubing between the thumb and forefinger, a channel is made, through which the liquid flows in a stream that is under complete control. The burette is held in an upright position, by means of a brass wire clamp, as shown in Fig. 29B.

FIG. 29B.



Filtration.—This is usually performed by means of a filter, in the form of a cone, set in a glass funnel. A circular piece of paper is twice folded, as shown in Fig. 30; then fitted to the funnel. It should be of such a size that it will not reach quite to the rim of the funnel. Unless otherwise stated, the filter should be wet with water and the water allowed to run through before the liquid to be filtered is poured on. If only a small quantity of liquid is to be filtered, the funnel may be set in the test-tube, Fig. 31, a strip of paper being inserted between the funnel and the tube to allow free passage of air.

FIG. 30.

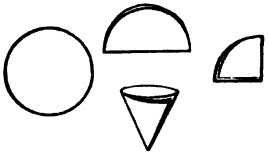


FIG. 31.

A large quantity of liquid should be filtered into a beaker, the funnel being supported by the ring or clamp. The lower

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end of the funnel must touch the beaker so that the liquid runs down its side gently, and loss by splashing is avoided. The liquid which runs through is called the *filtrate*. If not perfectly clear, it may be passed again through the same filter. The arrangement and manipulation above described is shown in Fig. 32. A mixture of chalk and water will suffice for an experiment on filtration.

Fig. 32.

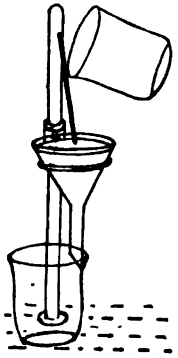

Fig. 34.

Washing Bottle, Fig. 33.—This is used for washing material in filters, dishes, etc. It may be easily made from an ordinary bottle and two pieces of glass tube passing through a doubly perforated stopper. By blowing in the shorter tube the water will be thrown out in a stream, and can be directed to any point.

Fig. 33.

Boring Corks.—Corks may be bored by the use of a small round file, Fig. 34. The cork is pierced by it, and the hole so made enlarged, by filing. Cork borers, Fig. 35, metal tubes of graduated sizes, with cutting edge, are convenient. A stout wire serves as a handle, and also to push out the

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If, instead of a closed tube, a jet is to be made, the tube should be drawn out to a narrow neck, Fig. 47, which is to be cut with a file at the point *a*.

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BENDING GLASS TUBES.—Tubes are bent by using the flaring flame obtained by means of the wing-top, Fig. 3, page 10, or in the flame of the ordinary gas jet. This permits several inches of the tube to be heated, and a regular curve, Fig. 38, to be made, instead of the sharp angle, Fig. 39. The tube should be held in the upper part of the flame, and slowly rotated so as to make the heating regular. When it begins to soften, it is removed from the flame and gently bent to the desired shape. It is then allowed to rest on the asbestos mat until cold.

FIG. 38.

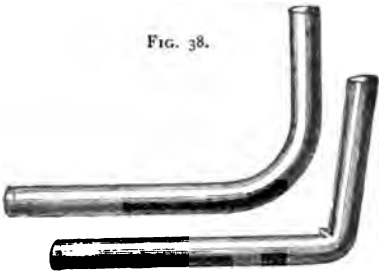
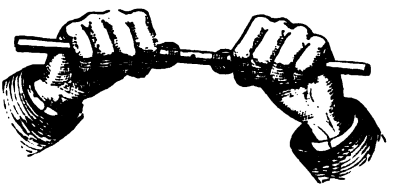


FIG. 39.

FIG. 40.



CUTTING GLASS TUBES AND RODS.—These are cut by making a deep scratch with a triangular file, at the point at which the division is to be made, and then bringing pressure to bear by the thumb and fingers of each hand in the manner shown in Fig. 40.

EVAPORATION OF LIQUIDS.—This is generally

performed in a porcelain dish, held in the ring as shown, Fig. 41.

If the evaporation is to be carried to dryness, that is, till all the liquid has been driven off, it is safest to rest the dish on a vessel containing boiling water, Fig. 42. Such an arrangement constitutes what is known as a *water bath*—properly a *steam bath.* It should be used whenever there is danger of overheating the contents of the dish.

Fig. 41.

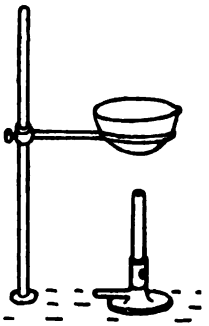
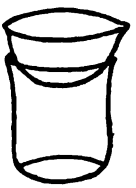


Fig. 42.



If only part of the liquid is to be evaporated, the boiling may take place in a beaker, as shown in Fig. 14, page 13. Partial evaporation is often resorted to as a means of "concentrating" a solution.

If the vapor from the liquid is collected and condensed, the process is called *distillation.* The condensed liquid is called the *distillate.*

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GENERAL CHEMICAL PRINCIPLES.

SOLUTION.

Exp. 1.—Measure 10 cc. of water into a test-tube, and add half a gram of powdered potassium chlorate. Holding the test-tube between the thumb and finger of one hand, tap it gently at the lower end, with the other hand; the potassium chlorate will gradually disappear. Preserve this liquid. The action is called *solution;* the body is said to *dissolve.* The following are additional examples:—

Exp. 2.—Put half a gram of copper into 10 cc. of water. It will produce a blue solution.

Exp. 3.—Add a very small quantity of potassium permanganate—a piece not larger than a pin's head—to 10 cc. of water; an intense purple solution will be formed.

Bodies that will not dissolve are said to be *insoluble.*

Exp. 4.—Put a small fragment of common rosin in about the same quantity of water as before. No solution will take place. Rosin is said, therefore, to be insoluble in water.

Exp. 5.—Take the test-tube used in the first experiment, add about one gram more potassium chlorate, and it will be found that not all will dissolve, even after prolonged shaking. The limit of solubility has been reached, and the solu-

tion is said to be *saturated.* Arrange the test-tube in the clamp stand, as shown in Fig. 12, and boil the liquid; all the potassium chlorate will dissolve. About two more grams should be added, and the boiling continued until this has been dissolved. This is the usual but not invariable influence of heat on the solubility of solids in liquids, namely to increase it.

Withdraw the heat and notice that as the tube cools, the solvent power of the water diminishes and the potassium chlorate separates in the form of small, glittering plates. When the liquid is cold, there is still some potassium chlorate in solution, only that portion separating which was rendered soluble by the heat.

Exp. 6.—Put a small quantity of flour—about as much as will be held on the end of a knife-blade—in a test-tube with some water, and shake vigorously. The flour will become thoroughly mixed with the water, and will not settle completely, even after standing some time. This condition is known as *suspension.* It must not be confounded with solution, in which a *clear* liquid is obtained.

Exp. 7.—Place in a test-tube a small piece of rosin, about the size of a pea, and pour over it 5 cc. of alcohol. In a short time some of it will dissolve, forming a clear solution. Now pour the alcoholic solution into a beaker half full of water; a turbid liquid is the result. The rosin

is insoluble in the mixture of alcohol and water, and separates from it in a finely divided condition. It is no longer in *solution* but in *suspension*. When in this manner a solid is formed by the action of one liquid upon another it is called a *precipitate*. The following are examples :—

Exp. 8.—To a solution made by adding a little barium nitrate to some water in a test-tube, add one drop of sulphuric acid. A white precipitate is formed.

Exp. 9.—Make, in separate test-tubes, solutions of lead nitrate and potassium chromate and mix them. A yellow precipitate is formed.

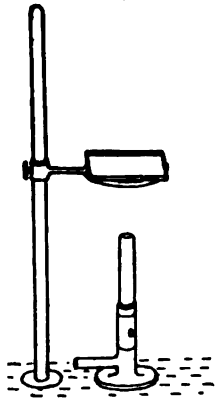
For convenience and rapidity in working, many chemical substances are kept in liquid form. Such bodies as are solid or gaseous at ordinary temperatures are used in solution, preferably in water, if soluble in that liquid. The liquids most frequently used are kept ready prepared on the laboratory table, and are collectively known as *reagents*. Of these sulphuric, nitric and acetic acids, and alcohol are naturally liquids; ammonium hydroxide and hydrochloric acid are solutions of gases in water; and most of the others are solutions of solids in water. Only those of the last class are now prepared in the laboratory, the various acids, alcohol and ammonium hydroxide being best made on the large scale. The process of making the ordinary reagent solutions is very simple. Some of the solid to be dis-

solved is put into the bottle, which is then nearly filled with water, the stopper is put in, and the bottle is shaken from time to time. Usually about one part of solid to twenty parts of water, by weight, is sufficient. In cases of sparingly soluble bodies, still less may be required.

PHYSICAL AND CHEMICAL CHANGE.

Exp. 10.—Counterpoise a small piece of steel, *e. g.*, a pen, accurately on the scale, then magnetize it by rubbing one end with a magnet. Bring the pen in contact with another pen which has not been touched with the magnet; the two will be mutually attracted. Place the magnetized pen again on the balance; the weight will be found to be unchanged, therefore nothing has been added to the steel, nor taken from it. This is an example of a *physical change.*

FIG. 43.



Exp. 11.—Take a piece of thin sheet iron, about two inches square, and bend up the edges all round, so as to make a shallow dish. Counterpoise it, accurately, on the balance by means of fine sand. Then weigh out on it one gram of that form of finely divided iron, known as reduced or Quevenne's iron. Place this on the ring-stand Fig. 43, and heat nearly to redness with the

Bunsen burner. In a few moments the powdered iron will begin to glow, and combustion will extend entirely through it. Allow it to cool; note that the color has changed to a reddish-brown; transfer the dish and contents, without loss, to the scale, and observe that a marked increase in weight has occurred.

In this experiment a change has been produced which affects the composition of the substance; something taken from the air has united with the iron, as is shown by the increase in weight. Changes of this kind are *chemical changes.*

Exp. 12.—Take about one gram of pure, dry zinc oxide, and spread it over a piece of sheet iron about two inches square. The dish used in the preceding experiment will answer. Rest this on the wire triangle and heat nearly to redness with the burner. The oxide turns bright yellow. Allow it to cool, and it will be found that the powder is again in every respect the same as before. If the piece of iron, with the zinc oxide on it, be weighed before and after the experiment, it will be found to have neither lost nor gained. This is a physical change.

Exp. 13.—Place at the bottom of an ignition-tube a fragment of sulphur, about the size of a cherry-stone. Heat cautiously with the flame turned low. It may be necessary to shut off the air supply entirely, to prevent the flame "running back." The sulphur will melt, and on further

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heating be converted into a red vapor, which will condense upon the upper portion of the tube either in drops or as a solid. When the tube cools, the whole mass will return to the solid condition. The sulphur has passed from the solid to liquid and liquid to gas, then by cooling back to the solid state.

The conversion of a solid into vapor and condensation again into solid form is known as *sublimation*. The condensed sulphur, deposited in the upper part of the tube, is termed a *sublimate*. These are physical changes.

Exp. 14.—Place in a small ignition tube a small quantity of mercuric iodide and heat gently, holding the tube above and not in the flame. The red powder soon begins to turn yellow. When all of it has apparently changed, allow it to cool and stir it with a glass rod. It will return to its original color wherever the rod touches it. These changes are physical, being unaccompanied by any change in composition.

Exp. 15.—Make a mixture of about equal parts of powdered sulphur and finely powdered copper. Note that however intimate the mixture may be, the particles of sulphur and copper can be distinguished on careful examination. If a portion of the mixture be stirred in water, the heavier copper will settle, and some of the lighter sulphur will float.

Place some of the *dry* mixture in the bottom

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of an ignition tube and heat strongly. The copper will glow and be converted into a bluish, brittle mass, from which neither copper nor sulphur can be extracted by water. Chemical combination has occurred and a compound has been formed.

CONDITIONS AFFECTING CHEMICAL ACTION.

The essential condition affecting chemical action is *intimate* contact of the bodies concerned. This is commonly expressed by saying that chemical affinity acts only over *insensible* distances. The significance of the phrase and the contrast between *chemical* and *physical* action are shown in the following experiments:—

Exp. 16.—Place on a white card a small quantity of iron filings, and move a magnet to and fro in contact with the under surface of the card. It will be seen that the influence of the magnet will extend through the card and affect the iron filings.

Exp. 17.—Put into a test-tube a fragment of sodium chloride and a fragment of silver nitrate, add water, and a white precipitate of silver chloride will form. If this is exposed to a bright light, best to direct sunlight, it will become violet-colored. That the light is the cause of the change, can be easily proved by preparing a second tube in the same manner and wrapping it in a piece of dark cloth so that the light cannot reach

it. It will not change in color. In this case the light which produces the effect proceeds from the sun and acts over a very great distance.

Exp. 18.—Make a tube as shown in Fig. 44 by folding a piece of sheet iron, six inches long by one inch wide, around a tube, about one-half an inch in diameter; then, holding it by the longer leg, heat the curved part in the flame until quite red-hot; withdraw it and insert quickly the tipped end of a match, taking care not to touch the metal. The heat will act over the intervening space and inflame the match.

FIG. 44.

Exp. 19. — Powder separately small quantities—about one-half gram each—of potassium ferrocyanide and ferric alum, taking care to clean and dry the mortar and pestle between each operation. Mix these powders carefully, on a piece of paper, with a spatula, avoiding pressure. No visible effect is produced unless the materials are moist, in which case the experiment must be repeated with perfectly dry substances. Return the mixture to the mortar, and rub it with the pestle; a blue color will at once appear, due to the chemical action brought about by close contact. Although the substances were in contact when mixed on the paper, they were not sufficiently close to allow the chemical action to occur.

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Exp. 20.—Copper sulphate and potassium ferrocyanide should be tried in the same manner, taking care, as before, to *wash* and *dry* the mortar and pestle between each powdering. A red color will appear when the bodies are rubbed together.

It is obvious that any condition that facilitates contact of bodies will favor chemical action. The degree of fineness of division will have much influence in this respect.

Exp. 21.—Iron, when heated in air, forms *iron oxide.*

(*a*) Heat to redness in the burner flame a piece of bright iron wire; only a superficial coating of iron oxide is formed.

(*b*) Place about one gram of Quevenne's iron on a piece of asbestos and apply a light. The iron will unite with the oxygen of the air, being in time entirely converted into iron oxide.

(*c*) Put into an ignition-tube about a gram of ferrous oxalate and heat nearly to redness. Iron in a very finely divided state remains in the tube. If the tube be now inverted, the iron will take fire spontaneously as it falls through the air.

The varying degrees of susceptibility of iron are thus seen to depend upon the fineness of the powder; the finest permitting the most thorough and rapid contact with the air.

Solution is a condition in which the particles of bodies are separated and diffused among

those of the dissolving liquid. The degree of division is far greater than that existing in the finest powder obtainable. Hence, solution is a most active promoter of chemical change. For this reason, as already stated, the majority of reagents are used in solution.

Exp. 22.—Prepare the dry mixture of ferric alum and potassium ferrocyanide, as directed in Exp. 19, but instead of rubbing it in a mortar throw it into some water in a test-tube. A copious blue precipitate is at once formed.

Exp. 23.—Copper sulphate and potassium ferrocyanide should be treated in the same manner.

Exp. 24.—Powder about one gram each of alum and sodium acid carbonate; mix them thoroughly when dry, and note that no action occurs. Divide the mixture into two equal parts and place each part in a *dry* test-tube. To one tube add some strong alcohol; to the other some water. It will be seen that although a liquid is present in both cases, no perceptible action takes place in the tube containing the alcohol, since the materials are not dissolved and are therefore not brought in closer contact. In the tube containing the water a chemical action attended by effervescence at once occurs, because the bodies present are soluble in water.

It is the fact of the solution, and not the

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mere presence of a liquid, that determines the action.

Solution may be either physical or chemical. In the first case the body diffuses through the particles of the dissolving liquid without change in composition, and often retains most of its properties, such as color, taste and odor. When the liquid is removed by evaporation or otherwise, the body is recovered, unchanged in composition.

Exp. 25.—Take a fragment of copper sulphate and dissolve it in about 10 cc. of water; the solution will be aided by heating. Place the liquid in an evaporating dish on the water bath, Fig. 42, page 22, and evaporate to dryness. The body will be recovered unchanged.

Exp. 26.—Add a drop of sulphuric acid to 5 cc. of water in a test-tube, then small quantities of magnesia, shaking between each addition, until no more dissolves. Filter the liquid, and evaporate it to dryness in a porcelain basin. A white solid having a very bitter taste is left. This is magnesium sulphate, formed by the action of the sulphuric acid on the magnesia. The magnesia dissolves in the acid, because it is acted upon chemically, it does not form a simple solution.

Operations like the above occur frequently in the laboratory and although the products are called solutions, it is necessary that the student

should appreciate the proper use of the word, thus:—

Exp. 27.—If a small fragment of copper be put into a few drops of nitric acid, a violent action will occur, red fumes will be evolved and the copper will disappear. The blue liquid which is formed is often called a solution of copper but is really a solution of copper nitrate, this body having been produced by the action of the acid and then dissolved in the liquid present.

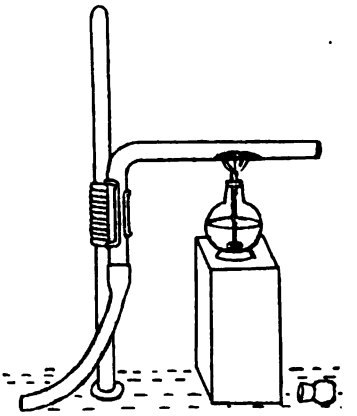
DECOMPOSITION—ELEMENTS AND COMPOUNDS.

Exp. 28.—Place about a gram of mercuric oxide in an ignition-tube and apply strong heat. A sublimate of mercury is formed and a gas given off in which bodies burn more vividly than in ordinary air. A match is lighted, then blown out so as to leave a spark on the end, and this is inserted in the tube. The spark will at once brighten and generally the wood will burst into flame, burning more brilliantly than in air. Two substances have been formed from the mercuric oxide, the mercury which collects on the side of the tube and the gas, oxygen, which escapes.

The result is called *decomposition*, the mercuric oxide being decomposed by the heat. With the means at present at our command neither the mercury nor the oxygen can be further decomposed.

Exp. 29.—Take a piece of glass tube six inches long and one-half inch internal diameter and bend an inch at one end at a right angle. Disconnect the Bunsen burner and attach the rubber tube to the shorter limb of the glass tube. Then mount the glass tube in the clamp so that the longer portion of it is horizontal, Fig. 45. Place at the middle of the horizontal limb a small amount of copper carbonate. This can easily be done by taking a strip of stout paper or thin sheet-iron, narrow enough to enter the tube, placing the heap of copper carbonate on the end of the strip, passing it into the tube and inverting it at the proper point. Heat the copper carbonate by means of a spirit lamp or burner from an adjacent table.

FIG. 45.



Copper carbonate consists of copper, carbon and oxygen. By heat it is decomposed into two substances—one a black solid which contains all the copper and part of the oxygen, and the other a colorless gas which contains the rest of the oxygen and all of the carbon. It is

evident that neither of these products represents the limit of decomposition. To demonstrate this, so far as the solid residue is concerned, turn on the current of gas, so that it may pass through the glass tube. To prevent escape of unburnt gas, light it at the other end of the tube, and turn it down so that the flame is about half an inch long. Apply heat to the residue and small particles of copper will appear. Care must be taken not to melt the tube. When the copper is observed, withdraw the heat, allow the gas to flow and continue to burn until the tube is cold, after which turn off the gas, remove the tube and empty out the metal. Copper has been set free by the decomposition of the copper oxide, the oxygen having been taken up by the coal gas. The copper cannot be further decomposed by any known means.

Exp. 30.—Into an ignition-tube put sufficient copper oxalate to make a layer about half an inch deep, and, holding the tube by the forceps or clamp, apply a strong heat. The copper oxalate contains copper, carbon and oxygen; by the heat the carbon and oxygen are driven off, and copper remains in the tube. If the substance be weighed before and after heating, it will be found to have lost much weight, proving that something has been driven off from it.

Exp. 31.—Heat gently in a dry test-tube a gram of ammonium dichromate. The mass will

soon become very hot, nitrogen and steam will be given off, and chromic oxide remain as a light green powder.

Bodies formed by the union of elements under the influence of definite chemical attraction are called *compounds*. Mere association of one substance with another is not sufficient to constitute a compound.

Mixtures exhibit the properties of the bodies they contain, but compounds almost always possess properties peculiar to them and not exhibited by their elements.

Exp. 32.—Place in an ignition-tube a small globule of mercury, and add a crystal of iodine; heat the mixture gently; a bright scarlet compound will be formed, part of it subliming as a yellow mass. Note that the fluid, silver-like mercury and the blue-black crystalline iodine combine chemically to produce a body which has no resemblance to either. It has also physiological and other properties different from either of the ingredients, being, among other qualities, much more poisonous.

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EXERCISES IN THE CHEMISTRY OF THE MORE IMPORTANT ELEMENTS AND COMPOUNDS.

HYDROGEN.

Exp. 33.—Half fill with water a wide-mouth vessel—bottle or tumbler—for which a piece of glass or card-board is provided as a cover, Fig. 46. Drop on the water a piece of sodium not larger than a cherry stone, and at once place on the cover. The sodium will melt, be driven over the surface of the water with a slightly hissing noise and gradually disappear. The action is due to the expulsion of a portion of the hydrogen of the water by sodium. (R.)*

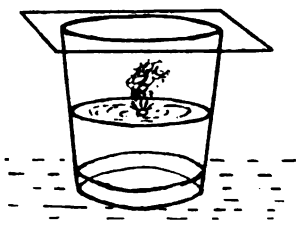

FIG. 46.

Exp. 34.—Into a test-tube one quarter full of water, put a piece of bright magnesium ribbon. No action will occur. Heat nearly to boiling. Bubbles of gas will accumulate on the surface of the magnesium. These consist of hydrogen expelled by the action of the magnesium. (R.) It will be seen that the effect is much less rapid

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* Whenever this symbol appears, it is intended that the student shall write out the reaction involved.

than in the case of the sodium, and almost entirely ceases when the heat is withdrawn.

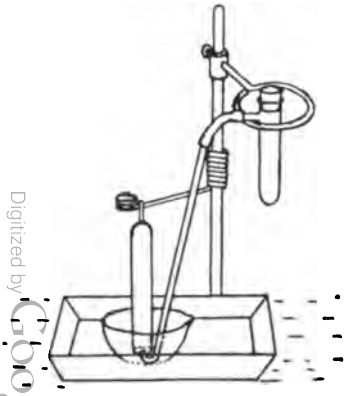
Exp. 35.—Repeat the last experiment, using a fragment of zinc instead of magnesium. No action will occur, even when the water is heated. Withdraw the heat and add a few drops of hydrochloric acid; effervescence immediately occurs. The interaction here is between the zinc and the hydrochloric acid. (R.)

FIG. 47.

Exp. 36.—Into a test-tube containing about ten cc. of sodium hydroxide solution, put a piece of aluminum, and warm gently. Hydrogen will be evolved. (R.)

To observe the properties of hydrogen, it will be necessary to collect it.

FIG. 48.



Exp. 37.—Fill a porcelain evaporating dish with water and place it in a flat glass basin; fill the gas receiver, Fig. 47, to the brim with water, cover it with a small piece of stout paper, pressed down so as to touch the rim all

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around. Invert the receiver into the water in the dish and support it by the helix, as shown in Fig. 48, so that the mouth of the tube is below the level of the water, and just over the deepest part of the dish.

By means of a rubber tube, attach to the side-neck test-tube, a glass tube, so arranged that one end dips in the water in the porcelain basin under the mouth of the gas receiver. Place in the side-neck tube some pieces of sheet zinc with 10 cc. of water, and not more than 2 cc. of hydrochloric acid. An immediate evolution of gas will occur, which will displace the water in the gas receiver. Any overflow from the dish will be caught by the shallow basin. When the receiver is full of gas, lift it, holding the mouth downward, and apply a light. A slight explosion may occur. Refill the receiver with water as before, and replace it in the dish so as to collect a second portion of the gas. Withdraw this, keeping the mouth downward, and apply a light. It will burn quietly. The combustion is due to the combination of the hydrogen with the oxygen of the air. (R.)

When the burning takes place rapidly, the water is converted into highly expanded steam, and an explosion occurs. If the receiver be only half filled with water, covered with paper, inverted as before and gas passed in, a mixture of hydrogen with air will be obtained. If the tube be raised out of the water and a light

applied to the mouth, combustion will occur at once, throughout the entire volume, with explosion.

ACIDS, BASES AND SALTS.

Exp. 38.—Put one drop each of sulphuric, nitric and hydrochloric acids, in separate test-tubes and fill each tube with water. Note the sour taste of either or all of these liquids. To each tube add a small piece of litmus paper and note that it is instantly reddened. Save these dilute acids for the following experiments.

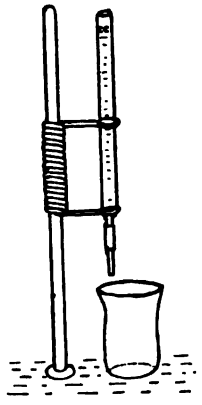
Exp. 39.—Put a portion of one of the dilute acids, prepared in the above experiment, with the litmus paper, in a test-tube, and add magnesia in small portions until the litmus paper turns blue. By the action of the magnesia, the acid is *neutralized* and a *salt* is formed. Bodies which neutralize acids in this manner, are called *bases.*

Exp. 40.—Take another portion of one of the dilute acids, drop into it a piece of phenolphthalein paper; the color will remain unchanged, but if magnesia be added as before, the paper will acquire a red tint, when the acid is neutralized. Here again, the color change indicates the neutralization of the acid. If the solution be poured off from the excess of magnesia, and a drop of acid be added to it, the paper will again become colorless.

Exp. 41.—Into separate test-tubes put small amounts of copper, copper oxide and copper carbonate, and add to each a few drops of nitric acid. Note that a greenish-blue solution of the salt, copper nitrate, is formed in all three cases. It is thus seen that a salt may be obtained by the action of an acid on either the metal, its oxide or carbonate. There are also other methods of making salts.

Exp. 42.—Put into a beaker 5 cc. of sodium hydroxide solution and drop in a piece of litmus paper. Prepare dilute hydrochloric acid by adding 5 cc. of strong acid to 20 cc. of water. These are best put into a flask or bottle and mixed by shaking.

FIG. 49.



Fill the burette, Fig. 49, with this liquid, place underneath it the beaker containing the sodium hydroxide and litmus paper and run in the acid, drop by drop, until the litmus paper is just reddened. Transfer the liquid from the beaker to the evaporating dish and evaporate it to dryness on the water bath. Note that the residue is common salt, which can be recognized by its taste. If perfectly dry, it will not affect neutral litmus paper. The product of this action is called a *salt*.

POTASSIUM.

Exp. 43.—Fill a tumbler half full of water, drop on the surface a small piece of potassium and cover the tumbler with a piece of glass, as directed in Exp. 33. The potassium will act upon the water, driving out hydrogen, which takes fire, and the vapor of the potassium produced at the same time gives a characteristic violet flame. (R.) After the potassium has disappeared, note the soapy feel of the liquid, due to the formation of an alkali; also note that red litmus is at once turned blue by it.

Exp. 44.—Place a piece of potassium nitrate, as large as a cherry stone, in not more than one cc. of water. Add to it a solution of tartaric acid, made in the same proportion. Pour the liquid into a watch-glass and stir briskly with a glass rod. Small crystals of potassium acid tartrate will appear. (R.) This reaction constitutes a test for potassium, but is only applicable to strong solutions.

SODIUM.

Examine a piece of sodium, note the softness, the lustre, and rapid tarnishing of a freshly cut surface.

Exp. 45.—Perform Exp. 43, using a small piece of sodium instead of potassium. The sodium will melt and act energetically on the water, but

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the escaping hydrogen will not inflame unless the sodium be delayed in movement, as when it touches the glass. Any flame produced will be yellow, from the vapor of sodium. The liquid produced has a soapy feel and an action on litmus paper like that produced in the potassium experiment.

SILVER.

Exp. 46.—Heat to dull redness a thin piece of silver foil by holding it, in the forceps, a short distance above the Bunsen flame. Note that, after cooling, the surface of the metal is still bright, no oxidation having taken place.

Exp. 47.—Place in a test-tube a small piece of silver, add a few cc. of water, drop in a small quantity of precipitated sulphur, and boil. A coating of silver sulphide is formed. Rinse off the silver with water and heat it to redness. The sulphide will be decomposed (reduced) and the metal will again become bright. Many of the silver compounds are reduced in this way by simple heating.

Exp. 48.—Prepare a solution of silver nitrate, by dissolving in nitric acid, in a test-tube, the piece of silver which has been used in the above experiments, using gentle heat if necessary. (R.) When all the silver is dissolved, dilute with about 5 cc. of water. Use this dilute solution for the experiments mentioned below. It will

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contain some excess of nitric acid which will not interfere, except in Experiment 52.

Exp. 49.—Place in a test-tube a few drops of silver nitrate solution, and add a few drops of a solution of ferrous sulphate. Heat the liquid and note the deposit of metallic silver which is formed on the sides of the tube. Ferric sulphate and ferric nitrate are formed at the same time. (R.) Many other reducing agents will reduce silver compounds in the same way.

Exp. 50.—Place in a test-tube a piece of copper foil and add a few drops of silver nitrate. The silver will be driven out of combination by the copper, copper nitrate being formed. (R.)

Exp. 51.—To a small quantity of water in a test-tube add a drop or two of silver nitrate solution and hydrochloric acid. A white curdy precipitate of silver chloride is formed, which remains undissolved on the addition of a few drops of nitric acid. (R.) This reaction furnishes a convenient test for chlorides.

Exp. 52.—For the following experiments the solution of silver nitrate must be neutral, that is, it must not contain any free acid. For this purpose either dissolve solid silver nitrate in distilled water, or shake finely powdered chalk with the remainder of the solution obtained in Experiment 48.

Exp. 53.—Place in a test-tube a drop or two of

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silver nitrate solution; add some water and a few drops of potassium chromate solution. A red precipitate of silver chromate is formed. (R.)

Exp. 54.—Place in a test-tube a few drops each of solutions of sodium chloride and potassium chromate. To the mixture add a drop of silver nitrate; the precipitate is white and consists wholly of silver chloride. (R.) Continue the addition of silver nitrate. After all the chlorine has combined and precipitated with the silver (not before) the silver will react with the potassium chromate to form silver chromate. The latter reaction will become evident by the red color of the precipitate. (R.)

These reactions are used in the quantitative estimation of chlorine in chlorides.

CHLORINE.

Exp. 55.—Put in a large test-tube a piece of potassium chlorate not larger than twice the size of a pin's head and pour upon it about one-half cc. of hydrochloric acid. A greenish-yellow gas is given off, which is a mixture of chlorine with chlorine tetroxide. (R.) When sufficient of the gas has been liberated to be recognized by its color (this is best seen against a white background) the tube should be half filled with water in order to stop the action, and obtain a solution of the gas. Preserve this solution (chlorine water) for the succeeding experiments. The gas should

not be inhaled, since it is highly irritating to the respiratory tract.

Exp. 56.—Make a solution of cochineal by boiling a few of the insects in some water in a test-tube, to which a drop of sodium hydroxide has been added to aid solution. Add some of the prepared chlorine solution; the color will be discharged.

Exp. 57.—Place in a test-tube filled with water a single drop of hydrogen sulphide. Add some of the chlorine water; the odor will be destroyed. (R.)

BROMINE.

Exp. 58.—In a test-tube, place a pinch of manganese dioxide, a minute crystal of potassium bromide, and one-half cc. of water; add five or ten drops of sulphuric acid and warm gently. Ruddy fumes of bromide will be liberated. Cautiously note the odor, which is similar to that of chlorine. (R.)

IODINE.

Exp. 59—Place in a test-tube, some manganese dioxide, a minute crystal of potassium iodide and a little water. Add five or ten drops of sulphuric acid and heat the mixture. Iodine will be liberated as a heavy violet vapor, which will condense in the upper part of the tube as a black crystalline sublimate. (R.)

Exp. 60.—Heat in a clean, dry test-tube, a minute crystal of iodine. The violet vapors of iodine can in this way be more readily observed.

Exp. 61.—Allow the test-tube in the above experiment to cool, and add about three cc. of water. Note that it becomes but slightly colored, very little solution taking place. Now add a crystal of potassium iodide, and shake gently. The iodine quickly dissolves, forming a dark brown liquid.

Exp. 62.—Put in a test-tube a fragment of iodine and a few cc. of alcohol. The iodine dissolves, forming tincture of iodine. Repeat this experiment, using chloroform instead of alcohol, and note that the solution is violet-colored. The cause of the difference in the color of these two solutions is not known.

Exp. 63.—Make some starch solution, as follows: Fill a test-tube half full of water and bring it to boiling. Take a piece of starch about the size of a cherry stone, put it in another tube and shake it up with some cold water. Add several drops of the milky liquid to the boiling water; allow the solution to cool, divide into two parts and put into one a single drop of an iodine solution. An intense blue color is developed. The production of this color is utilized as a test both for iodine and for starch.

Exp. 64.—Into the other tube containing starch solution add a small piece of potassium iodide; no color is produced. Add a single drop of the chlorine water obtained in Experiment 56; the color is developed at once. Iodine is liberated by the chlorine. (R.) It is thus seen that starch is a test for free iodine, and not for iodine in combination.

Of the three elements, chlorine, bromine and iodine, chlorine, under ordinary conditions, has the highest and iodine the least affinity. These facts are shown in the following experiments:—

Exp. 65.—Place in a test-tube about 1 cc. of chloroform and pour upon it about 5 cc. of chlorine water. Drop into the tube a crystal of potassium bromide, insert a cork and shake briskly. The chlorine will act upon the potassium bromide, driving out the bromine by its superior affinity. (R.)

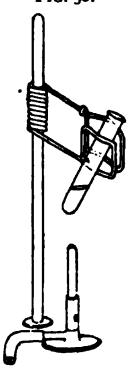
The chloroform takes up the bromine, forming a yellow or brown solution. Carefully pour off the watery liquid, as far as possible, without losing chloroform, and wash the latter by filling the tube with water and pouring off. Add 5 cc. of water and a piece of potassium iodide and shake briskly, as before. The bromine which was retained by the chloroform will drive the iodine out of combination. (R.) The iodine set free will dissolve in the chloroform, producing a

violet solution, which is most satisfactorily observed by pouring off the watery liquid, filling the test-tube with water and shaking, so as to extract the excess of bromine. In this experiment, chlorine has been made to drive out of combination bromine, which in its turn drives out iodine.

HYDROGEN CHLORIDE.

Exp. 66.—Place a test-tube in the clamp, as shown in Fig. 50; put into it one cc. of strong sulphuric acid and a few fragments of sodium chloride—common salt. Effervescence will take place, and a gas having a pungent, disagreeable odor will escape. It is hydrogen chloride, hydrochloric acid. (R.)

FIG. 50.



Exp. 67.—Hold over the mouth of the tube in which the hydrochloric acid is being liberated the wet stopper of the ammonium hydroxide bottle; dense white fumes of ammonium chloride are formed. (R.) The production of these fumes constitutes a test for either of these bodies.

Exp. 68.—Moisten a slip of blue litmus paper and hold it over the mouth of the tube containing hydrochloric acid. It turns red, showing the presence of an acid.

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Exp. 69.—The liquid called hydrochloric acid is a solution of hydrogen chloride in water; when the liquid is heated the hydrogen chloride is driven off. The escape of this gas may be shown by heating a few drops of the liquid in a test-tube held in the clamp and testing the vapors by the methods above given.

Exp. 70.—Put a drop of hydrochloric acid in some water in a test-tube, and add a few drops of silver nitrate solution. A white curdy precipitate is formed. This is silver chloride. (R.) Add to the liquid several drops of nitric acid and note that the precipitate remains undissolved.

The production, with a solution of silver nitrate, of a white precipitate, insoluble in nitric acid, constitutes a test for hydrochloric acid or for any other soluble chloride.

HYDROGEN BROMIDE.

Exp. 71.—Place in a test-tube a few cc. of strong orthophosphoric acid, drop in a few crystals of potassium bromide and heat to boiling. A gas escapes having an odor similar to that of hydrogen chloride. (R.)

NOTE.—Sulphuric acid cannot be used here in place of phosphoric acid, since hydrogen bromide is decomposed when heated with sulphuric acid.

Exp. 72.—Test the gas with moist blue litmus paper: it will turn the paper red.

Exp. 73.—Hold the wet stopper of the ammonium hydroxide bottle to the mouth of the tube; white fumes of ammonium bromide are formed. (R.)

Exp. 74.—Make a solution of potassium bromide by dissolving a small crystal in a few cc. of water. Add to this some silver nitrate solution. A yellowish-white precipitate of silver bromide is formed, which is insoluble in nitric acid. (R).

HYDROGEN IODIDE.

Exp. 75.—Prepare hydrogen iodide in the same way as hydrogen bromide was made, substituting potassium iodide for the potassium bromide. Note the action of the gas on litmus paper; also the fumes with ammonium hydroxide. (R.)

Exp. 76.—Dissolve a small fragment of potassium iodide in water and add silver nitrate solution. A yellow precipitate of silver iodide is formed. (R.) The precipitate is not dissolved by nitric acid. This reaction constitutes a test for iodides.

HYDROGEN FLUORIDE.

Exp. 77.—Make an intimate mixture, by rubbing in a mortar, of small amounts of powdered calcium fluoride and sodium acid sulphate. Introduce the mixture into an ignition-tube, and heat the end of the tube in the flame. Vapors

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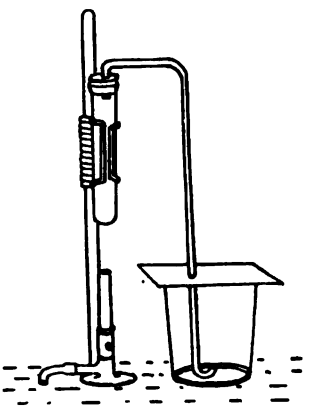
of hydrogen fluoride will rise, and corrode the glass. Allow the tube to cool, break off the end, wash out the water, dry, and note the etched surface.

OXYGEN.

Exp. 78.—Heat, in an ignition-tube, about a gram of potassium chlorate. After some time it will melt, and on further heating effervesce. Plunge into the escaping gas, the glowing end of a match stick; it will be re-inflamed. The gas is oxygen.

FIG. 51.



A more convenient method of obtaining and collecting oxygen is shown in the next experiment.

Exp. 79.—Make an intimate mixture, by rubbing in a mortar, of two grams of potassium chlorate and half a gram of manganese dioxide. Introduce the mixture into a *dry* test-tube, for which there has been provided a perforated cork and delivery tube. The test-tube is supported in the clamp so that the delivery tube may dip to the bottom of a beaker or tumbler (Fig. 51). A gentle heat is applied to the mixture, and in a

few seconds the gas in the receiving vessel is tested by a glowing splinter of wood, as in the previous experiment. It will be re-inflamed. (R.)

The manganese dioxide is not changed, but causes the potassium chloride to decompose more easily and regularly. The exact manner in which the effect is produced is not known.

Exp. 80.—To establish the fact that in the above experiment the potassium chlorate has been changed to chloride, allow the test-tube to cool, add water, shake, filter and treat the filtrate with silver nitrate and nitric acid—a few drops of each. A white precipitate of silver chloride is formed. (R.) A crystal of potassium chlorate dissolved in water, tested in the same way, will not give this precipitate.

OZONE.

Exp. 81.—Put a pinch of barium dioxide into a small beaker and add a few drops of strong sulphuric acid (without water). Barium sulphate and water are formed and oxygen liberated, some of the oxygen in the form of ozone. (R.) Note the characteristic odor of ozone. Save the contents of the beaker.

Exp. 82.—Prepare a starch solution (see Exp. 63), and to this liquid add a small quantity of potassium iodide; the mixture should be colorless. Dip a strip of filter paper in the mixture and hold it in the beaker containing the ozone.

Ordinary oxygen does not decompose potassium iodide, and therefore produces no color on the paper; but ozone liberates iodine, which colors the starch blue. (R.)

Exp. 83.—Prepare ozone by placing a small quantity of potassium permanganate in a test-tube and adding a few drops of sulphuric acid (without water) as in Exp. 81. The reaction results in the formation of permanganic anhydride, potassium sulphate and water. (R.) The permanganic anhydride decomposes, yielding ozone and lower oxides of manganese. Test for ozone with potassium-iodide-starch paper, as before.

Exp. 84.—Cover the test-tube used in Exp. 83 with a clean silver coin; in a few minutes the surface of the coin will be tarnished. This discoloration is due to the formation of silver oxide. (R.)

N. B. The discoloration of silver exposed to the air is not due to ozone but to the presence of sulphides in the atmosphere.

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HYDROGEN DIOXIDE.

Nascent oxygen in the presence of water forms hydrogen dioxide. It is also formed when dilute acids act upon certain highly oxygenized bodies.

Exp. 85.—Mix in a test-tube twenty drops of sulphuric acid and three cc. of water. Mix in a porcelain basin a small piece of barium dioxide

with sufficient water to form a thick paste. By means of a glass rod, add small portions of this paste to the test-tube containing the dilute sulphuric acid; hydrogen dioxide will be formed. To prevent its decomposition, the tube containing the sulphuric acid should be kept cool by immersion in a beaker of cold water. (R.)

Exp. 86.—To demonstrate the bleaching power of hydrogen dioxide, make a solution of extract of logwood by boiling it in water for a short time. Allow this to cool, and render it alkaline by a few drops of ammonium hydroxide. Pour one drop of this into a test-tube containing water, and add twenty drops of commercial solution of hydrogen dioxide. (The solution of hydrogen dioxide prepared in Exp. 85 will probably not be strong enough to show the effect.) The color is promptly discharged, only a slight yellow tint remaining. To prove that this action is not due to any acid present, add several drops of ammonium hydroxide, which should give no color.

Exp. 87.—Place in a test-tube about three cc. of commercial solution of hydrogen dioxide, and add some manganese dioxide. An evolution of oxygen will occur.

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OXIDATION AND REDUCTION.

Exp. 88.—Ignite some zinc turnings by holding them in the flame by the aid of the tongs. Zinc oxide, a white powder, is formed. (R.)

Exp. 89.—Burn a small amount of Quevenne's iron by heating it on a piece of thin sheet-iron; ferric oxide is formed. (R.)

Exp. 90.—Make a solution of ferrous sulphate by dissolving a small crystal in a few cc. of water in a test-tube; add to this, sodium hydroxide solution; a greenish precipitate of ferrous hydroxide is formed. (R.) Pour the contents of the tube into a beaker half full of water, and then pour this dilute liquid from one beaker to another several times, so as to expose it freely to the air. Oxygen is taken up and ferric hydroxide formed, the precipitate becoming red. (R.)

Exp. 91.—Put a small quantity of powdered lead sulphide (the crystalline natural form, known as galena) in a test-tube, and add a few drops each of water and of nitric acid. Lead sulphate is formed by direct addition of oxygen. (R.)

Exp. 92.—Put into an ignition-tube a small quantity of mercurous iodide; add a fragment of iodine and a few drops of alcohol, and shake. Red mercuric iodide is formed. (R.)

Exp. 93.—Place a small quantity of silver oxide in an ignition-tube, and heat gently. The oxygen is set free and the silver is reduced, forming a bright mass. (R.)

Exp. 94.—Arrange the apparatus as shown in Fig. 45, Exp. 29. Introduce into the middle of the horizontal limb of the glass tube a small

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amount of powdered ferric oxide. Turn on the coal gas, light it at the mouth of the tube, turning it down so as to give a flame about one-half inch long. Heat the ferric oxide, by means of a spirit lamp, until it turns black. Withdraw the heat, and continue the current of gas until the tube is cold, then turn off the gas, and empty the contents of the tube upon a piece of sheet iron. The black powder is *reduced* or *Quevenne's Iron*, obtained by reduction of the ferric oxide by the hydrogen of the coal gas. (R.)

Exp. 95.—Make silver chloride, by mixing small amounts of silver nitrate and sodium chloride in a test-tube, with water, and expose it to direct sunlight for a few moments. It becomes of a violet color. Silver chloride is partly reduced by the light, some of the chlorine being liberated.

Exp. 96.—Add a fragment of potassium chromate to a mixture of five drops of sulphuric acid with the same quantity of water. Chromic acid is formed. (R.) Add a small quantity of powdered zinc; the red liquid becomes rapidly green. The chromic acid present has been reduced by the aid of the zinc and the sulphuric acid into chromic sulphate; water and zinc sulphate are formed at the same time. (R.)

Exp. 97.—Place in a test-tube a fragment of sodium sulphite; add some water and ten drops of sulphuric acid. Sulphurous acid is formed.

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(R.) In another test-tube dissolve a *small* amount of potassium permanganate in water, using heat if necessary.

Mix these solutions and the color of the permanganate will be instantly discharged. The sulphurous acid has been oxidized to sulphuric and the permanganate reduced to potassium sulphate and manganous sulphate. (R.)

SULPHUR.

Exp. 98.—Place in an ignition-tube a small piece of sulphur and heat gently. The sulphur will first melt and then be converted into a gas, which will condense in the upper and cooler portion of the tube. This process is used in making ordinary flowers of sulphur (sublimed sulphur).

Exp. 99.—Place in a test-tube a small quantity of powdered sulphur, and add to it a few cc. of water and a small quantity of milk of lime (slaked lime made into a paste with water). Heat for a few minutes, filter, and add to this solution (which now contains a higher sulphide of calcium) hydrochloric acid, drop by drop, until a white precipitate appears. This precipitate is finely divided sulphur, known as *precipitated* sulphur.

Exp. 100.—Heat a piece of sulphur in an ignition-tube until it melts and vapor of sulphur appears, then pour the molten mass into cold water. When cool, it will be a brownish-yellow and elastic material, an allotropic form (plastic sulphur).

Exp. 101.—Prepare copper sulphide as follows: Place in an ignition-tube a small piece of sulphur, and over it, several strips of copper foil, ½ by 2 inches, rolled into a helix. Place the copper helix so that it will lie midway in the tube, and not in contact with the sulphur. Heat the tube first in the vicinity of the copper, and then at the lower end, so that the vapor of sulphur shall pass over the red-hot copper. A chemical combination, with an evolution of heat, occurs. Copper sulphide is formed. (R.) Contrast the bluish, brittle copper sulphide with the original copper.

Exp. 102.—Place a small quantity of flowers of sulphur, with some water, in a test-tube. Shake the mixture, add a bright piece of silver, and boil. The coin blackens from the production of silver sulphide. (R.) This is the so-called oxidized silver used in jewelry.

Exp. 103.—Place some sulphur at the bottom of a narrow ignition-tube and introduce an equal quantity of Quevenne's iron. Heat first the portion of the tube which contains the iron, then that which contains the sulphur. Combination occurs and ferrous sulphide is formed. (R.) When the tube is cold, break it and place a portion of the cinder-like mass in a test-tube. Add several drops of hydrochloric acid. Hydrogen sulphide will be evolved as a colorless, odorous gas. (R.) If the gas be not immediately evolved, heat the tube. Cover the test-tube.

Dip a filter paper into a solution of copper sulphate, and hold it in the test-tube. It turns brownish from the formation of copper sulphide. (R.)

Repeat the experiment with a piece of paper dipped in a solution of lead nitrate. A black stain, due to lead sulphide, is formed.

Exp. 104.—Prepare separate solutions of copper sulphate, zinc sulphate, lead nitrate and ferric chloride. Add to each small portions of hydrogen sulphide solution and note the changes. (R.)

N. B. In the fourth test-tube the precipitate is free sulphur, ferrous chloride being produced at the same time.

Exp. 105.—Take 2 cc. of hydrogen sulphide solution and add to it one drop of sodium hydroxide. Sodium hydrosulphide is formed. (R.) Add to the liquid one drop of ferric chloride solution. Ferrous sulphide, sodium chloride, hydrogen sulphide and sulphur, are formed. (R.)

SULPHUROUS ACID.

Exp. 106.—Cover the end of a stick with powdered sulphur, and ignite the sulphur by holding it in the flame of the burner. Hold the burning sulphur in the mouth of a test-tube containing about 3 cc. of water. Sulphurous anhydride, a colorless, irritating gas, is formed. (R.) Cover the mouth of the test-tube and shake

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vigorously until the anhydride is absorbed by the water; sulphurous acid is formed. (R.)

Exp. 107.—Burn several portions of sulphur in the same manner indicated in the previous experiment, and obtain a stronger solution of sulphurous acid. Divide the contents of the test-tube into two portions.

Exp. 108.—Test the bleaching power of sulphurous anhydride by adding to the contents of one of the tubes a drop or two of a very weak alkaline solution of logwood extract, or a rose petal.

Exp. 109.—To the remaining portion of the sulphurous acid add a solution of hydrogen sulphide until a white precipitate of free sulphur appears. Smell the contents of the tube and note that the odor of hydrogen sulphide has been removed by the sulphurous acid. (R.)

Exp. 110.—Place a fragment of sodium sulphite in a test-tube; add several drops of sodium hydroxide solution and 2 cc. of silver nitrate solution. Boil. The sodium sulphite will reduce the silver nitrate and form a mirror of silver on the inner surface of the tube. Sodium sulphate, sodium nitrate, water and silver are formed. (R.)

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SULPHURIC ACID (HYDROGEN SULPHATE).

Exp. 111.—Place in a test-tube a pinch of sodium sulphate; add from five to ten drops of nitric acid and heat gently. Sodium sulphate

will be formed and nitric oxide liberated. (R.) Preserve this solution for Exp. 115.

Exp. 112.—Place in a test-tube about one cc. of sulphuric acid and add an equal bulk of water. Hold the tube at the bottom so that the hand encircles it, and note the development of heat, due to the strong affinity of the acid for water.

Exp. 113.—Place in a test-tube a small quantity of sugar and add a few drops of sulphuric acid. A black carbonaceous mass is formed. This change is due to the abstraction from the sugar of the hydrogen and oxygen which exist in it in the proportion to form water. Carbon, the remaining element, is set free.

Exp. 114.—Marks made on paper, with dilute sulphuric acid and dried over the flame, will show the same charring action on organic matter.

Exp. 115.—Place in a test-tube a drop of sulphuric acid and a few cc. of water. Add to this several drops of barium chloride solution. A white precipitate of barium sulphate is formed. (R.) This is one of the most insoluble bodies known. The production of this white precipitate with barium salts is utilized as a test for sulphates. Other bodies (carbonates, phosphates, etc.), also form precipitates with barium salts, but these are all soluble in acid. The addition of acid is, therefore, necessary to make the test complete. Take the solution prepared in Exp. 111, and

test it for a sulphate by the addition of barium chloride and hydrochloric acid.

CALCIUM.

Exp. 116.—To ten drops calcium chloride solution add one drop of sulphuric acid. Calcium sulphate is precipitated. As this is slightly soluble, no water should be added, otherwise, the precipitate may not appear. (R)

Exp. 117.—To a drop of calcium chloride solution add some water and a drop of solution of sodium phosphate. A white precipitate of sodium phosphate is formed. (R.) Add several drops of hydrochloric acid to the mixture and note that the calcium phosphate dissolves. This action explains the use of hydrochloric acid in dissolving from bones the mineral matter, which is largely calcium phosphate.

Exp. 118.—To a few drops of calcium chloride solution add a drop or two of sodium carbonate solution. A white precipitate of calcium carbonate is formed. (R.) This is the so-called *precipitated* chalk.

Exp. 119.—The ordinary lime-light is produced by heating a piece of calcium oxide (lime) to incandescence in the flame of an oxy-hydrogen blow-pipe. The same effect may be produced by heating a very thin film of the oxide in the Bunsen flame. This film may be prepared by dipping a piece of filter paper in a very dilute

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solution of calcium chloride, drying, and burning the paper to an ash, by which the calcium chloride is converted into calcium oxide. If the ash be held in the margin of the flame, it will become luminous. Many other infusible opaque oxides, *e. g.*, magnesia and baryta, show the same effect.

Exp. 120.—Prepare calcium hydroxide by mixing quicklime with ten times its volume of water, and allow it to stand until the lime has become converted into a soft white powder. (R.) Shake the mixture and filter some of it rapidly into a test-tube. The clear filtrate is calcium hydroxide solution; most of the calcium hydroxide, however, remains undissolved. Blow into the filtrate through a glass tube. The expired carbon dioxide will form a white precipitate of calcium carbonate. (R.)

BARIUM.

Exp. 121.—Add to one drop of the solution of barium nitrate, a little water, and then a drop of sulphuric acid. A white precipitate of barium sulphate is formed, which is insoluble in acids. (R.) Add to the mixture several drops of hydrochloric and several drops of nitric acid; note that the precipitate remains unaffected.

Barium compounds impart to flame a bright green color.

Exp. 122.—Dip the end of a piece of iron wire in a solution of barium chloride and hold it in

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the flame of a Bunsen burner. Note the green color which is produced.

STRONTIUM.

Exp. 123.—A solution of any strontium compound treated as in Exp. 122 will impart a crimson-red color to flame.

The other reactions of barium and strontium are similar to those of calcium.

LEAD.

Exp. 124.—Place in a test-tube a few grains of red lead-oxide; add a few drops of water and of nitric acid. Heat the mixture The red-lead will be decomposed, lead dioxide being formed, which remains as a brown powder, and lead nitrate, which dissolves. (R.)

Lead is but little affected by hydrochloric acid or sulphuric acid, but is dissolved by hot nitric acid.

Exp. 125.—Put in a test-tube a small piece of lead and add to it some water and five or ten drops of nitric acid. Heat the liquid. Lead nitrate is formed, which dissolves. (R.) The nitric acid acts first as an oxidizing agent.

Exp. 126.—To a single drop of lead nitrate solution add a little water and a fragment of potassium iodide. Yellow crystalline lead iodide is precipitated. (R.) Heat the liquid and the

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precipitate will be found to dissolve. Allow it to cool, and the crystals will be again deposited.

Exp. 127.—To a few drops of lead nitrate solution add a drop of sulphuric acid. A heavy white precipitate of lead sulphate will be formed. (R.)

Exp. 128.—To a solution of lead nitrate add a drop or two of ammonium carbonate; lead carbonate, mixed with some lead hydroxide, is precipitated. (R.)

Exp. 129.—To a drop of lead nitrate solution, add a little water and then a drop of the solution of potassium chromate. A yellow precipitate of lead chromate is formed. (R.)

Exp. 130.—Place in a test-tube a small piece of zinc, and cover it with lead nitrate solution. The lead will be driven out of combination by the zinc, and will be deposited as a crystalline mass. (R.)

Exp. 131.—Put a single drop of lead nitrate solution into a test-tube, and fill the tube nearly full with water; then pour out all but a few drops of the liquid, and refill with water. Add to this water—now containing a very small quantity of lead nitrate—a drop of ammonium sulphide. A brownish-black precipitate will be formed. This is lead sulphide. (R.) The formation of this dark precipitate by ammonium sulphide constitutes one of the delicate tests for

lead, but it alone is not sufficient to show the presence of lead, since copper, iron and some other metals give a similar color.

ZINC.

Exp. 132.—Hold in the Bunsen flame, by means of the tongs, a loose bunch of zinc turnings, and note the rapid combustion and production of zinc oxide. (R.)

Exp. 133.—Collect some of the zinc oxide produced in Exp. 132, and note that when cold, it is white, but that if it be heated on a piece of sheet iron, it assumes a yellow tint, becoming white again on cooling.

Exp. 134.—Place a piece of zinc with some water in a test-tube, and add a few drops of sulphuric acid. The effervescence is due to the evolution of hydrogen, zinc sulphate being formed. (R.)

Exp. 135.—Place in a test-tube, a small quantity of zinc oxide, with some water, and add a few drops of sulphuric acid. The zinc oxide dissolves, but no evolution of gas occurs. (R.)

N. B. If the zinc oxide made by burning zinc in the air be used for this experiment, care must be taken that it contains no unburnt zinc. It is best to use pure, recently calcined zinc oxide.

Dissolve a small quantity of zinc sulphate in water, and divide the solution into three parts.

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Exp. 136.—To one part add a few drops of sodium hydroxide; a gelatinous precipitate of zinc is formed. (R.) The precipitate dissolves in excess of the sodium hydroxide.

Exp. 137.—To the second portion add a solution of sodium carbonate; zinc carbonate mixed with some zinc hydroxide is formed. (R.)

Exp. 138.—To the third portion, which should be made alkaline by the addition of a little ammonium hydroxide, add a few drops of hydrogen sulphide solution. White zinc sulphide is formed. (R.)

MAGNESIUM.

Exp. 139.—Hold in the Bunsen flame, by means of the tongs, a piece of magnesium about four inches long. Note the rapid combustion and the production of white magnesium oxide. (R.)

Exp. 140.—Place a piece of magnesium ribbon (about one inch long) in a wide-mouth test-tube. Add water and a few drops of sulphuric acid. The effervescence which occurs is due to the evolution of hydrogen, displaced by magnesium, magnesium sulphate being formed.

Exp. 141.—To a few drops of a solution of magnesium sulphate, add sodium hydroxide solution; a white precipitate of magnesium hydroxide is formed. (R.)

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Exp. 142.—To a few drops of solution of magnesium sulphate solution, add a solution of sodium carbonate; a white precipitate of magnesium carbonate is formed. (R.)

Exp. 143.—To a few drops of a solution of magnesium sulphate, add several drops each of ammonium hydroxide and sodium phosphate solution, then stir the mixture with a glass rod. A crystalline precipitate of ammonium-magnesium phosphate is formed. (R.)

COPPER.

Exp. 144.—Put in a test-tube a small piece of copper, cover with water and add a few drops of hydrochloric acid. No effect takes place even when the liquid is boiled.

Exp. 145.—Pour off the liquid from the above and add water and sulphuric acid. The result is similar.

Exp. 146.—Pour off the sulphuric acid from the copper, and add a few drops of nitric acid; the metal is attacked at once. Copper nitrate, water and nitric oxide are formed. (R.)

Exp. 147.—Heat a piece of copper foil for several minutes in the air. It will take up oxygen and form black copper oxide. (R.) Divide the piece of oxidized copper into two portions; place each in a test-tube and add to one water and hydrochloric acid, and to the other water

and sulphuric acid. Heat gently; the liquids become faintly colored, showing that the copper oxide is being dissolved. (R.) It will be noted that as the action proceeds the surface of the metal becomes cleaner. It is this solvent action upon metallic oxides that renders dilute acids serviceable in cleaning tarnished surfaces. The tarnish is generally a film of oxide, and this is taken up by the dilute acid while the metal proper is but little affected.

Exp. 148.—Place in a test-tube five or ten drops of copper sulphate solution, add a little water and about ten drops of sodium hydroxide solution. A blue precipitate of copper hydroxide is formed. (R.) Boil the liquid; the copper hydroxide is decomposed, copper monoxide and water are formed. (R.)

Exp. 149.—To five or ten drops of copper sulphate solution, add a few drops of a strong solution of sodium sulphite, then about ten drops of sodium hydroxide solution, and boil as before. In this case a salmon-red precipitate is formed. Sodium sulphite is a reducing agent and takes oxygen away from the mixture, forming copper suboxide—cuprous oxide. (R.)

Exp. 150.—Place in a test-tube five or ten drops of copper sulphate, some water, and a few drops of ammonium hydroxide. A pale bluish-white precipitate is formed, which contains copper hydroxide; now add more ammonium hydroxide;

the precipitate dissolves, forming a clear blue solution. The production of this blue color is utilized as a test for copper.

Exp. 151.—Put a small piece of clean bright iron in a test-tube, and pour over it a little copper sulphate solution; the copper is driven out of combination by the iron, and deposited as a bright red film. (R.)

Exp. 152.—To a few drops of copper sulphate solution in a test-tube, add some sodium carbonate; a bluish precipitate of copper carbonate is formed. (R.)

Exp. 153.—To a drop of copper sulphate solution in some water in a test-tube, add one drop of ammonium sulphide; black copper sulphide is precipitated. (R.)

Exp. 154.—Mix a few drops of copper sulphate with a few drops of potassium ferrocyanide; brown copper ferrocyanide is precipitated. This reaction is utilized as a test.

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MERCURY.

Exp. 155.—Place in a clean, dry test-tube, a small drop of mercury and heat it, holding the tube a short distance above the flame, which should be turned down rather low. The mercury boils and the vapor condenses in liquid globules in the upper and cooler part of the test-tube.

Mercury, like copper, is not dissolved by dilute acids, except nitric acid.

Exp. 156.—Into the test-tube containing mercury used in Exp. 155 place about ten drops of water and five drops of hydrochloric acid. Pour off the acid from the mercury and add diluted sulphuric acid. Note that in each case the mercury remains unaffected.

Exp. 157.—Pour off the sulphuric acid from the mercury in Exp. 156, and dissolve the latter *completely* in nitric acid by the aid of heat.. (R.) The nitric acid acts as an oxidizing agent. The solution should be preserved for Exp. 160.

Exp. 158.—Place in a test-tube a small quantity of mercury; add some water and then nitric acid, drop by drop, until bubbles of gas begin to come off from the mercury. Let the acid act for several minutes. No heat is to be applied. In this case mercurous nitrate is formed. (R.) The reaction is similar to that in Exp. 157.

Exp. 159.—Add to the solution of mercurous nitrate (obtained in Exp. 158) a drop of hydrochloric acid. A white precipitate of mercurous chloride is formed. (R.)

Exp. 160.—To the solution of mercuric nitrate (obtained in Exp. 157) add a drop of hydrochloric acid. No precipitate is formed. Mercuric chloride being soluble in water cannot be formed by

precipitation. It is usually made by sublimation, as shown in the next experiment.

Exp. 161.—Mix intimately small quantities of mercuric sulphate and sodium chloride. Heat the mixture in an ignition-tube and note the formation of a white sublimate on the cooler portion of the tube. This is mercuric chloride (corrosive sublimate). (R.)

Exp. 162.—Place in a test-tube a few drops of a solution of mercuric chloride and add one drop of solution of potassium iodide. A scarlet precipitate of mercuric iodide is formed. (R.) Add several more drops of potassium iodide solution. The mercuric iodide will dissolve, forming a clear liquid.

Exp. 163—To a few drops of a solution of mercuric chloride, add a drop of sodium hydroxide solution. A yellow precipitate of mercuric hydroxide is formed. (R.)

Exp. 164.—Place in a test-tube a piece of copper; cover it with a solution of mercuric chloride and warm gently. The copper will become coated with a bright deposit of mercury. (R.) If the copper slip be removed, dried and heated in an ignition-tube, the mercury will sublime and form globules in the cooler portion of the tube. A few drops of hydrochloric acid added to the solution will cause the mercury to

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deposit on the copper more quickly. In this form the reaction is known as *Reinsch's test.*

ALUMINUM.

Aluminum is acted upon both by alkalies and acids, especially hydrochloric acid.

Exp. 165.—Place in a test-tube a small piece of aluminum, cover it with water and add a few drops of hydrochloric acid; hydrogen is evolved and aluminum chloride formed, which dissolves. (R.) Save the solution.

Exp. 166.—Place in a test-tube a small piece of aluminum and cover it with sodium hydroxide solution. Hydrogen is evolved and a compound formed called sodium aluminate. (R.)

Exp. 167.—To the solution of aluminum chloride obtained in Exp. 165, add a single drop of sodium hydroxide solution. A white precipitate of aluminum hydroxide is formed. (R.)

Exp. 168.—Make a solution of cochineal by boiling several of the insects in water. Add to this several drops each of aluminum chloride and ammonium hydroxide solutions. Aluminum hydroxide is precipitated and carries the coloring matter with it, leaving the liquid colorless. Precipitates of this kind are technically known as *lakes.*

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IRON.

Exp. 169.—Take a small portion of Quevenne's iron and heat it on a piece of thin sheet iron. It will take fire and burn, producing ferric oxide. (R.)

Exp. 170.—Place in a test-tube a small piece of iron wire, cover it with water and add several drops of hydrochloric acid. The iron is acted upon, ferrous chloride is formed, which produces a yellowish-green solution, and hydrogen escapes. (R.)

Exp. 171.—To the solution of ferrous chloride obtained in Exp. 170 add several drops of sodium hydroxide solution. A greenish precipitate of ferrous hydroxide is formed. (R.) Note that at the surface, where it is in contact with the air, the precipitate gradually becomes reddish. This is due to the absorption of oxygen and formation of a ferric compound.

Exp. 172.—Make a solution of ferrous chloride by treating a small piece of iron with excess of hydrochloric acid. Divide this in two parts; to one add hydrogen dioxide, and note that the liquid which was at first yellowish-green, turns bright yellow. This is due to the formation of ferric chloride. (R.) Save this liquid.

Exp. 173.—To the other portion of the ferrous chloride solution, add a few drops of nitric acid, and boil. Ferric chloride is formed. (R.)

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Exp. 174.—Add to the ferric chloride solution obtained in Exp. 172 sodium hydroxide, drop by drop, until a reddish-yellow precipitate of ferric hydroxide is formed. (R.)

Exp. 175.—Place some powdered zinc in a test-tube, and add a solution of ferric chloride. The zinc reduces the ferric chloride to ferrous chloride, the solution becoming nearly colorless. (R.)

Exp. 176.—To a solution of ferric chloride in a test-tube, add several drops of sodium carbonate solution and heat the liquid to boiling. A reddish precipitate is formed, which consists largely of ferric hydroxide along with a small amount of ferric carbonate.

CHROMIUM.

Exp. 177.—Make a mixture of about equal parts of potassium nitrate and chromic oxide. Place it on a piece of thin sheet iron and heat until the mass fuses and turns brownish-yellow. When the mass is cold, put a piece in water and note the yellow solution produced, due to potassium chromate, formed through the oxidation of the chromic oxide by the potassium nitrate. (Potassium nitrate acts as an oxidizing agent by breaking up into oxygen, potassium oxide and nitric oxide.)

Exp. 178.—To a part of the solution of potassium chromate, obtained in Exp. 174, add a drop

of lead nitrate. Yellow lead chromate is precipitated. (R.)

Exp. 179.—To another portion of the solution of potassium chromate add several drops of sulphuric acid, and then a reducing agent, such as sodium sulphite. The chromium is reduced from the hexad to the double-tetrad condition, green chromic sulphate being formed. (R.)

MANGANESE.

Potassium permanganate is made by methods of oxidation similar to that used for the preparation of potassium chromate.

Exp. 180.—Bend a piece of sheet iron so as to form a shallow dish; place in this a small quantity each of potassium or sodium hydroxide, potassium nitrate and manganese dioxide. Heat the mixture until it fuses and turns dark green. Allow it to cool and place it in a beaker of water. A green solution of potassium manganate is obtained.

Exp. 181.—Pour into a test-tube some of the clear green solution obtained in the last experiment, and add to it a drop or two of nitric acid. The solution will turn bright red, potassium permanganate being formed. (R.)

Exp. 182.—Dissolve a fragment of potassium permanganate in some water in a test-tube, and add to it several drops of sulphuric acid and a

drop or two of sodium sulphite solution. The color of the potassium permanganate is instantly discharged. The sodium sulphite has decomposed it by appropriating some oxygen and becoming sodium sulphate. (R.)

NICKEL.

Exp. 183.—Dissolve a small piece of nickel in nitric acid by aid of heat; nickelous nitrate is formed. (R.) Dilute the deep green liquid with several times its volume of water and divide it into three portions.

Exp. 184.—To one of these portions add ammonium hydroxide, a precipitate of nickelous hydroxide is first formed (R.), which, by the addition of excess of ammonium hydroxide, is redissolved, forming a blue solution.

Exp. 185.—To a second portion of the nickelous nitrate solution, add sufficient sodium hydroxide to render the liquid alkaline, and then add sodium hypochlorite. A black precipitate of nickelic hydroxide is formed. (R.)

Exp. 186.—To the third portion of the solution of nickelous nitrate, add sufficient sodium hydroxide to render it alkaline, and then add hydrogen sulphide. A black precipitate of nickelous sulphide is formed. (R.)

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COBALT.

Exp. 187.—Dissolve a small piece of cobalt in nitric acid; cobaltous nitrate is formed. (R.) Dilute the solution with several times its volume of water and divide it into three parts.

Exp. 188.—Add to one portion of the solution of cobalt nitrate, sodium hydroxide until a precipitate forms, and boil the liquid. A dirty-red cobaltous hydroxide is formed. (R.)

Exp. 189.—To a second portion of the solution of cobalt nitrate, add some alcohol and sodium hydroxide and boil. A brownish-red precipitate of cobaltic hydroxide is formed. (R.)

Exp. 190.—Add to the third portion of the solution sufficient sodium hydroxide to render it slightly alkaline, and then add hydrogen sulphide. Cobaltous sulphide is precipitated. (R).

NITROGEN.

NITROUS OXIDE.

Exp. 191.—In a dry test-tube to which is attached a cork and bent tube leading into a beaker or tumbler (Fig. 52), place several grams of ammonium nitrate and heat with the burner. The salt melts and gives off nitrous oxide and water. (R.) The gas, being heavier than air, collects in the receiving vessel.

Plunge into the gas, a stick with a glowing end; the wood will at once burst into flame,

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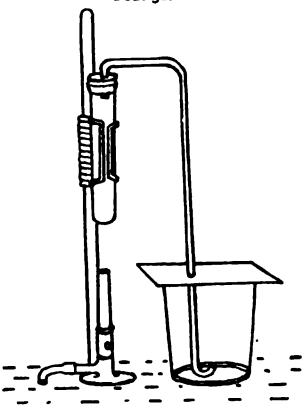
showing that combustion is supported more actively in nitrous oxide than in the air.

NITRIC OXIDE.

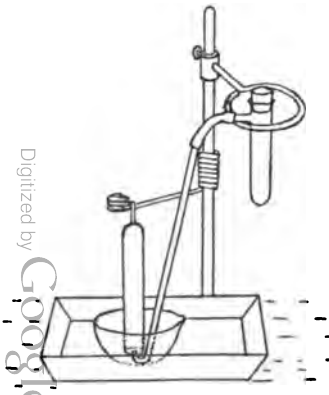
FIG. 52.



Exp. 192.—Fill the gas-receiver with water, cover the mouth with a piece of paper, invert it in a dish, as shown in Fig. 53.

Place in a test-tube provided with a cork and bent tube, or in the side-neck tube, some copper and cover it with a mixture of one volume of nitric acid with three volumes of water. Insert the cork with its tube and, if necessary, apply gentle heat. Copper nitrate, water and nitric oxide are formed. (R.) In a few moments, when all the air has been driven out by the escaping gas, thrust the end of the delivery tube under the receiver, so that the gas collects

FIG. 53.



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in it. When it is half filled remove the delivery tube.

Exp. 193.—Raise the receiver out of the water, so that air mixes with the gas; a brownish-red color is developed. This is due to the formation of nitrogen dioxide by union with the oxygen of the air. (R.)

NITRIC ACID.

Exp. 194.—Place in a test-tube a piece of zinc, cover it with water and add a few drops of nitric acid. Hydrogen is given off and zinc nitrate formed. (R.)

Exp. 195.—Place in a test-tube a small piece of tin-foil and add to it several drops of nitric acid. The tin is acted upon with energy, tin dioxide, water and nitric oxide being formed. (R.)

Exp. 196.—Place on a piece of thin sheet iron a small quantity of potassium nitrate until it fuses. Now sprinkle on this small pieces of paper or wood, and note how quickly combustion takes place. The carbon of the paper combines with the oxygen of the potassium nitrate, forming carbon dioxide.

AMINE and AMMONIUM COMPOUNDS.

Exp. 197.—Make a mixture of about equal parts of magnesium oxide and ammonium chloride. Place this in a test-tube and warm *gently.*

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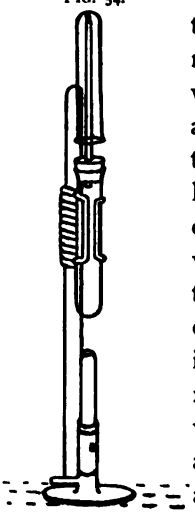
Cautiously smell the gas which is given off. It is amine (ammonia). (R.)

FIG. 54.



Exp. 198.—Insert in the test-tube containing ammonium chloride and magnesia, a cork provided with a straight glass tube and invert over it a second test-tube, as shown in Fig. 54. Apply heat to the mixture; amine will be evolved, and being lighter than air, will collect in the inverted test-tube. (R.) In a few minutes withdraw the upper tube, and keeping it inverted, so that the gas may not escape, place it mouth downward in a vessel of water. The amine will be at once absorbed, and the level of the water will rapidly rise to the top of the tube. A chemical action is supposed to take place here, the amine combining with water to form ammonium hydroxide. (R.)

Exp. 199.—Fill a tube with amine, and place it upside down in water, but when the latter has risen about one-fourth of an inch, cork the tube quickly and shake vigorously for a moment. A stronger solution of the ammonium hydroxide is obtained. The following properties should be noted. The solution is caustic; if the fingers be wetted by it, they will be found to have the

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peculiar soapy feel produced by caustic bodies. It is strongly alkaline; a piece of red litmus paper dropped in the solution is at once turned blue.

Exp. 200.—Take two watch glasses of equal size, place in one a drop of strong hydrochloric acid and in the other a drop of strong ammonium hydroxide solution. Invert one glass over the other. Dense white fumes are formed which finally deposit as a white solid; this is ammonium chloride, formed by the union of the hydrochloric acid with amine, which is always given off from strong solutions of ammonium hydroxide. (R.)

Exp. 201.—Heat a small quantity of ammonium chloride on a piece of sheet iron; it is quickly converted into vapor.

Exp. 202.—Place in a small glass tube a piece of ordinary glue, a substance containing carbon, hydrogen, nitrogen and oxygen; and heat with the Bunsen burner. The glue is decomposed and a part of the nitrogen combines with hydrogen to form amine. To test for its presence moisten a piece of red litmus paper and place it over the tube; in a short time the amine will turn the paper blue.

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PHOSPHORUS.

PHOSPHATES.

Exp. 203.—Dissolve sodium phosphate in water and test separate portions of the solution, with solution of calcium chloride, copper sulphate and ferric chloride. Calcium phosphate (R.), copper phosphate (R.) and ferric phosphate (R.) will be formed. These precipitates are soluble in hydrochloric acid.

Exp. 204.—Place in a test-tube a small piece of glacial phosphoric acid (metaphosphoric acid) and add to it a solution of egg-albumin. The albumin is coagulated.

Exp. 205.—Mix in a test-tube a few drops of orthophosphoric acid and solution of egg-albumin. No coagulation will occur.

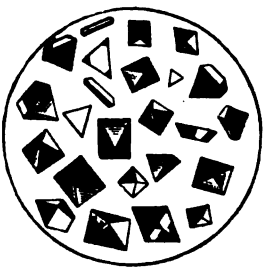
Exp. 206.—Place in a test-tube a small quantity of sodium hypophosphite, add water and a few drops of silver nitrate, and apply heat. A white precipitate, which soon changes to brownish black, is formed. This precipitate is silver, set free by the reducing action of the sodium hypophosphite, which is converted into phosphate.

Exp. 207.—Put a small quantity of sodium hypophosphite into an ignition-tube and heat it. A decomposition occurs by which phosphine, sodium pyrophosphate and water are formed. (R.)

ARSENICUM.

Exp. 208.—Place in an ignition-tube a pinch of arsenious oxide and heat gently. The oxide will sublime without decomposition and collect in the upper part of the test-tube in the form of small crystals which the microscope shows to be octahedral. (Fig. 55.)

FIG. 55.



Exp. 209.—Place in an ignition-tube a pinch of arsenious oxide, and on this, a layer, one-half inch in depth, of small pieces of charcoal (carbon). Heat the charcoal nearly to redness, then heat the lower portion of the tube so that the vapor of the arsenious oxide passes over the red-hot charcoal. The carbon will unite with the oxygen of the arsenious oxide, forming carbon monoxide, which escapes, and arsenicum which is deposited as a black metallic-looking ring, known as the arsenical mirror, in the upper part of the tube. (R.)

Exp. 210.—Make a solution of arsenious oxide by adding a very small amount to some water in a test-tube, and boiling for several minutes. Arsenious oxide is but slightly soluble in water, and when added to it appears to exert a repellant action which is almost characteristic. Save this solution.

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Exp. 211.—To a portion of the arsenious oxide solution, add a drop of hydrochloric acid, and then some hydrogen sulphide solution. A lemon-yellow precipitate of arsenious sulphide is formed. (R.) The hydrochloric acid does not enter into the reaction, but ensures the formation of the precipitate, since arsenious sulphide is less soluble in the acid than in pure water.

Exp. 212.—Reinsch's Test.—Place in a test-tube a strip of copper, cover it with water and add a few drops of hydrochloric acid. Boil the mixture. If the chemicals are pure no stain will appear on the copper. Now add a drop of the solution of arsenious oxide and boil. The arsenicum will be deposited upon the copper and form copper arsenide as a steel-gray film. The hydrochloric acid acts upon the arsenious oxide, forming arsenious chloride and water. (R.) The arsenious chloride reacts with the copper, copper chloride and arsenicum being formed. (R.) The arsenicum deposits on the copper and alloys with it. The copper slip is taken out of the liquid, dried by pressing between filter paper (the slip should not be rubbed, otherwise the film may be removed). It is then placed in an ignition-tube and heated. The arsenicum will be oxidized to arsenious oxide, which volatilizes and collects in the upper part of the tube in the form of octahedral crystals.

Exp. 213.—Repeat Exp. 212, adding, before

the boiling is begun, a small crystal of potassium chlorate, or other oxidizing agent. No deposit occurs upon the copper slip. Reinsch's test, therefore, cannot be used in the presence of strong oxidizing agents. With arsenates, the deposit on the copper can only be obtained by using large excess of hydrochloric acid.

ANTIMONY.

Exp. 214.—Place about ten drops of nitric acid, without water, in a test-tube, and add a pinch of powdered antimony. Warm gently. Oxidation takes place. Both antimonous and antimonic oxides are formed. (R.) The latter unites with water to produce metantimonic acid. (R.)

Exp. 215.—Prepare antimonous chloride by boiling a small amount of powdered antimony in a test-tube for a few minutes with a mixture of twenty drops of hydrochloric acid and five drops of nitric acid. (R.) Pour the liquid into a test-tube two-thirds full of water. A white precipitate of antimony oxychloride is formed. (R.) By continued action of water upon this, antimonous oxide is formed. Save the contents of this test-tube.

Exp. 216.—Collect on a filter some of the precipitate formed in Exp. 215. Wash it several times with water, and transfer it to a test-tube. Add to it a small quantity of cream of tartar (acid potassium tartrate) and 3 cc. of water, and boil

the mixture for a few minutes. Allow any undissolved precipitate to subside. The action in this experiment is the substitution of antimony monoxide (antimonyl), a monad radicle, for one atom of hydrogen of the acid tartrate. (R.) The product is tartar emetic.

Exp. 217.—Perform Reinsch's test as described under arsenicum, substituting for the arsenious oxide solution, the clear solution of tartar emetic obtained in the preceding experiment. The surface of the copper becomes coated with a dark stain, often having a somewhat bluish tint. This is a deposit of antimony. (R.) Pour the liquid out of the test-tube, wash the copper with a little water, dry it by pressing it between folds of porous paper, roll it up into small bulk and insert it into the closed end of an ignition-tube. On applying heat to the tube it will be found that the antimony does not volatilize as easily as the arsenical deposit does.

Exp. 218.—Introduce a few cc. of hydrogen sulphide solution into a test-tube and add a drop or two of the tartar emetic solution. An orange-red precipitate of antimonous sulphide is thrown down. (R.)

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BISMUTH.

Exp. 219.—Prepare bismuth nitrate by heating a fragment of bismuth with ten drops of strong nitric acid. (R.) Add this solution to a large

volume of water in a beaker. A white precipitate of bismuth oxynitrate (subnitrate) is formed by the action of the water on the normal nitrate. (R.)

Exp. 220.—Boettger's Test.—Boil together small quantities of bismuth oxynitrate and glucose in some water to which about ten drops of sodium hydroxide solution have been added. A heavy black precipitate, consisting in part of free bismuth, is formed.

Exp. 221.—Add hydrogen sulphide solution to the liquid obtained in Exp. 219. The white precipitate will become black, from the formation of bismuthous sulphide. (R.)

BORON.

Exp. 222.—Make a hot saturated solution of borax (sodium tetraborate) in two cc. of water; add ten drops of hydrochloric acid and allow the mixture to cool. Orthoboric acid is precipitated in white scales. (R.) Save the contents of the tube.

Exp. 223.—Decant the liquid from the precipitated boric acid formed in the last experiment, and add to the residue one cc. of alcohol, which will dissolve it. Dip a glass rod into the solution, hold it in the Bunsen flame and note the characteristic bright green mantle of the flame, due to boron.

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GOLD.

Exp. 224.—Place in a test-tube a small piece of gold, and add from two to four drops of nitric acid; no change occurs. Add five to ten drops of hydrochloric acid; free chlorine is formed which attacks the gold, forming auric chloride. This dissolves, forming a yellow solution, which should be saved. (R.)

Exp. 225.—To the solution of auric chloride, obtained in Exp. 224, add one cc. of strong solution of ferrous chloride, made by dissolving a small piece of iron in some hydrochloric acid. The gold is reduced to the metallic state and appears as a precipitate, ferric chloride being formed at the same time. (R.) If in appreciable amount, the gold may be collected upon a filter, and will, on rubbing, show a metallic lustre.

CARBON.

Exp. 226.—Place in an ignition-tube a piece of wood, and heat it to redness. The wood is decomposed; vapors are given off which will burn at the mouth of the tube, while carbon (wood charcoal) will be left in the tube.

Exp. 227.—Heat some fragments of bone in the same manner; vapors are also given off, but more disagreeable in odor than in the last experiment; carbon (animal charcoal), mixed with the mineral matter of the bone, remains.

Exp. 228.—Put two or three drops of hydrogen sulphide in a test-tube half filled with water; add about five grams of charcoal in small pieces, and cork the tube and shake it for a moment. In about fifteen minutes the hydrogen sulphide will be absorbed by the charcoal, and the odor will have disappeared.

Exp. 229.—Dissolve about 0.2 gram of quinine sulphate in 5 cc. of water by aid of a few drops of sulphuric acid. The peculiar bluish sheen known as fluorescence will be distinctly seen in this solution. If the liquid be shaken with a few grams of fresh animal charcoal and then filtered, the filtrate will not fluoresce, or only feebly, indicating the abstraction of quinine from the solution. The effect of the charcoal is also made evident by noting that the intensely bitter taste is removed entirely or in great part.

METHANE.

Exp. 230.—Mix intimately, by rubbing in a mortar, equal parts of sodium acetate and magnesia. Introduce this into an ignition-tube and apply strong heat. Methane will escape and may be ignited at the mouth of the tube. (R.) The flame will have a yellow tint due to the presence of a small quantity of sodium.

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CARBON DISULPHIDE.

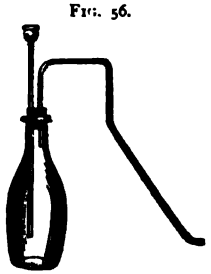
Exp. 231.—Place in a beaker a few drops of carbon disulphide; cover the mouth of the beaker with a piece of glass. In a few minutes remove it and insert into the beaker a small hot glass rod (it need not be red hot). The carbon disulphide vapor will inflame. Some of the sulphur may be deposited upon the sides of the beaker, but the greater portion, as well as the carbon, is oxidized. (R.)

Exp. 232.—Make a solution of sodium sulphide in a little water. To this add a small quantity of carbon disulphide, and warm gently (not to boiling). The liquid becomes of a light orange-red color due to the formation of sodium sulphocarbonate. (R.)

CARBON DIOXIDE.

Exp. 233.—Place in the gas bottle, provided with a funnel and delivery tube (Fig. 56), a few pieces of marble; cover them well with water, and insert the cork and tubes, allowing the delivery tube to reach to the bottom of a tumbler or beaker. Pour in through the funnel tube five cc. of hydrochloric acid. Carbon dioxide will be liberated and collect in the

FIG. 56.



receiving vessel by downward displacement. (R.) The mouth of the beaker should be covered with a piece of cardboard. Test for the gas by plunging into it a burning slip of wood; it will be extinguished.

Exp. 234.—Hold a piece of magnesium ribbon about four inches long in the forceps or helix; ignite the free end by holding it in the flame and then plunge it in the beaker containing carbon dioxide. The combustion will continue, magnesium oxide being formed, and fine particles of carbon be deposited on the sides of the vessel. (R.)

Exp. 235.—Arrange the gas-generating apparatus so that the evolved carbon dioxide may bubble through about five cc. of lime water (see Exp. 120) in a test-tube. If the evolution of gas has slackened, more hydrochloric acid should be added through the funnel tube. The lime water becomes rapidly turbid, from the formation of calcium carbonate. (R.) If the current of gas be continued for some minutes, the precipitate will disappear, being dissolved in the excess of carbon dioxide. Remove the delivery tube and boil the contents of the test-tube briskly for a few minutes. The excess of carbon dioxide will be driven off, and the calcium carbonate again precipitated.

Exp. 236.—Allow carbon dioxide to bubble

through a solution of litmus. A wine-red color is produced.

CARBON MONOXIDE.

Exp. 237.—Place in a wide test-tube a gram of potassium ferrocyanide and add two cc. of strong sulphuric acid. Hold the test-tube in the clamp as shown in Fig. 12, page 13, and heat carefully. The mixture will foam considerably, producing ferrous sulphate, potassium sulphate, ammonium sulphate, and carbon monoxide. (R.) The carbon monoxide may be ignited at the mouth of the tube. (R.)

Exp. 238.—Heat, in a wide ignition-tube, a mixture of about one gram each of magnesium carbonate and powdered zinc. The magnesium carbonate decomposes into magnesia and carbon dioxide. (R.) The latter is deoxidized by the zinc, furnishing carbon monoxide (R.), which may be ignited at the mouth of the tube.

SILICON.

Exp. 239.—Fold a piece of magnesium ribbon, about two inches long, into four thicknesses, insert it into the lower end of a *narrow* ignition-tube and heat strongly until the magnesium is fused into the glass. Magnesium silicide is formed wherever the magnesium is in contact with the glass. Allow the tube to cool, break it and put the fragments containing the magnesium into an evaporating dish, and cover them with

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strong sulphuric acid. Bubbles of hydrogen silicide are formed, which on coming in contact with the air inflame spontaneously with a crackling noise. (R.)

Exp. 240.—Add a few drops of soluble glass solution (solution of sodium orthosilicate) to some water in a test-tube, and then add a few few drops of calcium chloride solution. Calcium orthosilicate is formed. (R.)

Exp. 241.—Add one cc. of hydrochloric acid to a dilute solution of sodium orthosilicate in a test-tube. Orthosilicic acid is precipitated. (R.)

TIN.

For the following experiments care should be taken to use thin, pure, tin-foil.

Exp. 242.—Ignite a slip of tin-foil by holding it for a moment in the flame. The tin continues to burn for a short time after removal, forming a mixture of stannous and stannic oxides. (R.)

Exp. 243.—Treat a piece of foil with nitric acid, warming if necessary, until red fumes escape. Stannic oxide (anhydride) is formed (R.), which unites with water to form metastannic acid, a white powder. (R.)

Exp. 244.—Prepare a solution of stannous chloride by dissolving tin-foil in a small quantity of hydrochloric acid, warming if necessary. (R.)

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Dilute the solution and divide it into three portions for the next experiments.

Exp. 245.—To one portion of the solution of stannous chloride add hydrogen sulphide solution. A brownish precipitate of stannous sulphide is formed. (R.)

Exp. 246.—To another portion of the solution add a small amount of solution of mercuric chloride. Mercurous chloride is first formed as a white precipitate. (R.) This is gradually converted into a black precipitate of metallic mercury. (R.) The stannous chloride is converted by these reactions into stannic chloride.

Exp. 247.—To another portion of the solution of stannous chloride add a small piece of zinc. A black deposit of tin is produced on the surface of the zinc. (R.)

Exp. 248.—Prepare stannic chloride by dissolving tin in a mixture of ten drops of hydrochloric acid and five drops of nitric acid, warming, if necessary. (R.) When the action has ceased, add sodium hydroxide solution drop by drop until a precipitate is formed, which does not dissolve on shaking. This is stannic acid. (R.) Continue the addition of sodium hydroxide, and the precipitate will redissolve, owing to the formation of sodium stannate. (R.)

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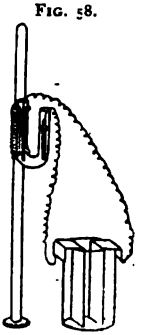
ELECTRICAL DECOMPOSITION.

A battery, such as shown in Fig. 57, will suffice for the following experiments: It consists of two plates, one of sheet copper, the other of amalgamated zinc, each six inches long by two inches wide, placed in a tumbler, nearly filled with a mixture of one part of sulphuric acid with ten parts of water. One end of each sheet is connected to a piece of No. 20 cotton-covered wire. The free ends of the wires are wound at least 12 times around sticks of dense carbon, two inches long and one-sixth inch in diameter. The bare wire should be wrapped very tightly, in order to ensure good electrical connection.

Fig. 57.

The decomposing cell which carries the liquid to be experimented upon, is made of a glass tube one-fourth inch calibre, five inches long, bent into U form, and held in the clamp. A carbon terminal (technically known as an electrode) is inserted into each limb of the tube. The liquid to be decomposed should be about half the quantity necessary to fill the cell. The whole arrangement is shown in Fig. 58. The terminal attached to the wire from the copper plate is called the positive pole or anode; that attached to the zinc

Fig. 58.



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is called the negative pole or cathode. The decomposing cell and electrodes must be thoroughly rinsed before each experiment.

Exp. 249.—Introduce into the decomposing cell a solution of mercuric iodide, prepared by adding to a saturated aqueous solution of mercuric chloride a strong solution of potassium iodide, drop by drop, until the red precipitate first formed is dissolved. Insert the electrodes and allow the current to pass for a few moments. The mercuric iodide is decomposed, the iodine at the anode forming a dense brownish solution, and the mercury appearing at the cathode, as a gray film which in time will collect into minute globules.

Exp. 250.—Place in the decomposing cell a strong solution of copper sulphate to which a small quantity of sulphuric acid has been added, and pass the current. A deposit of copper will form on the cathode, while the liquid around the anode will become strongly acid from the liberation of sulphuric anhydride and oxygen, the former immediately uniting with the water to produce sulphuric acid. The oxygen is produced in this case only in small amount and may be dissolved by the water.

Exp. 251—Mix in a beaker, two cc. of sulphuric acid with six cc. of water, and subject the liquid to the action of the battery, as before. Hydrogen is liberated at the cathode, escaping in bubbles,

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while oxygen and sulphuric anhydride are set free at the anode. The sulphuric anhydride unites with some of the water to re-form sulphuric acid.

When the element liberated at the cathode is of sufficient affinity to decompose water, it will not appear in visible form, but hydrogen will be evolved instead.

Exp. 252.—Subject to the action of the current a strong solution of common salt (sodium chloride), to which a drop or two of a solution of phenolphthalein has been added. Sodium will be liberated at the cathode, and will immediately decompose the water, liberating hydrogen and forming sodium hydroxide, which will redden the phenolphthalein. Chlorine will be set free at the anode, and may, if desired, be detected by adding a small amount of a solution of potassium iodide and starch, which will become blue.

Exp. 253.—Subject to the action of the current a strong solution of sodium sulphate which has been tinted with a little infusion of purple cabbage. The liquid at the anode will turn red from the presence of sulphuric acid, while that at the cathode will turn green from the presence of sodium hydroxide.

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Table of Symbols, Valencies, Electrical Relations and Atomic Weights of Elements Considered in this Work.

Positive, monad.			*Negative, monad.*		
Hydrogen,	H,	1.	Chlorine,	Cl,	35.5.
Potassium,	K,	39.	Bromine,	Br,	80.
Sodium,	Na,	23.	Iodine,	I,	127.
Silver,	Ag,	108.	Fluorine,	F,	19.

Positive, dyad.			*Negative, dyad.*		
Calcium,	Ca,	40.	Oxygen,	O,	16.
Barium,	Ba,	137.	Sulphur,	S,	32.
Strontium,	Sr,	87.5.			
Lead,	Pb,	207.			
Zinc,	Zn,	65.5.			
Magnesium,	Mg,	24.			
Copper,	Cu,	63.5.			
Mercury,	Hg,	200.			

Positive, dyad and tetrad.			*Intermediate, triad or pentad.*		
Aluminum,	Al,	27.	Nitrogen,	N,	14.
Iron,	Fe,	56.	Phosphorus,	P,	31.
Manganese,	Mn,	55.	Arsenic,	As,	75.
Chromium,	Cr,	52.	Antimony,	Sb,	122.
Nickel,	Ni,	59.	Boron,	B,	11.
Cobalt,	Co,	59.	Bismuth,	Bi,	208.
			Gold,	Au,	197.

Positive, tetrad.		
Carbon,	C,	12.
Silicon,	Si,	28.
Tin,	Sn,	118.

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Equivalents for Some English and French Measures.

The unit of measure in the French system is the meter, which is approximately one ten-millionth of the distance from the Equator to the North Pole, and equals 39.368 American inches.

One meter equals { 1000 millimeters. / 100 centimeters. / 10 decimeters.

1 cubic decimeter = 1 litre = 1000 cubic centimeters.

The unit of weight is the gram, the weight of one cubic centimeter (cc.) of water at a standard temperature and pressure.

1000 grams = 1 kilogram.
100 " = 1 hectogram.
10 " = 1 decagram.

0.1 gram = 1 decigram.
0.01 " = 1 centigram.
0.001 " = 1 milligram.

1 gram	= 15.432 grains.	1 grain	= 0.0648 gram.
2 "	= 30.864 "	2 "	= 0.1296 "
3 "	= 46.297 "	3 "	= 0.1944 "
4 "	= 61.729 "	4 "	= 0.2592 "
5 "	= 77.161 "	5 "	= 0.3240 "

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INDEX.

Acid, 41.
Alkali, 43.
Aluminum, 75.
Amine, 82.
Ammonia, 82, 83.
Ammonium compounds, 82.
Anode, 98.
Antimony, 88.
Arsenical mirror, 86.
Arsenicum, 86.
Arsenious oxide, 86.
Atomic weights, 101.

Balance, 14.
Barium, 65.
Base, 41.
Battery, 98.
Beaker, 12.
Bending tubes, 21.
Bismuth, 89.
Boiling, 12, 13.
Borax, 90.
Boric acid, 90.
Boring corks, 19.
Boron, 90.
Bromine, 47.
Bunsen burner, 9.
Burette, 17.

Calcium, 64.
Carbon, 91.
——— dioxide, 93.
——— disulphide, 93.
——— monoxide, 95.
Cathode, 99.
Centimeter, 102.
Chalk, 64.
Charcoal, 91.
Chemical change, 27.
Chlorine, 46.
Chromium, 77.
Clamp, 13.
Cobalt, 80.
Compounds, 34-37.
Copper, 70.
Cork borers, 19.
Cutting tubes, 21.

Decomposing cell, 34.
Decomposition, 98.
Distillation, 22.
Drawing out tubes, 20.

Electrical decomposition, 98.
Electrode, 98.
Elements, 34.
Evaporating dish, 22.

Filtrate, 19.
Filtration, 18, 19.
Flame, 10.

Gas receiver, 39.
Gram, 91.
Gold, 102.

Helix, 15.
Hydrogen, 38.
Hydrochloric acid, 50.
Hydrogen bromide, 51.
——— chloride, 50.
——— dioxide, 55.
——— fluoride, 52.
——— iodide, 52.
Hypophosphite, 85.

Ignition tubes, 20.
Insoluble, 23.
Iodine, 47.
Iron, 76.

Lakes, 75.
Lead, 66.

Magnesium, 69.
Manganese, 78.
Measurement, 14.
Mercury, 72.
Methane, 92.
Mortar, 13.

Neutral, 41, 45.
Nickel, 79.
Nitric acid, 82.
——— oxide, 81.
Nitrogen, 80.
Nitrous oxide, 80.

Oxidation, 56.
Oxygen, 53.
Ozone, 54.

Pestle, 15
Phosphates, 85.
Phosphine, 85.
Phosphorus, 54.
Physical change, 60.
Pipette, 17.
Precipitate, 75.
Potassium, 43
Pouring, 14, 17
Pyrophosphate, 85.

Quevenne's iron, 58.

Reinsch's test, 87, 89
Reduction, 56

Salt, 41, 49
Saturated solution, 94.
Silicon, 95.
Silver, 44.
Sodium, 43.
Soluble, 93
—— glass, 96.
Solution, 92, 93
Spatula, 14
Spirit lamp, 11
Stannic salts, 96
Starch solution, 98
Steam bath, 20
Stannous, 96
Strontium, 66
Sub-nitrate, 98.
Sulphur, 59.
Sulphuric acid, 80.
Sulphurous acid, 81.
Suspension, 94.
Symbols, 101.

Test tube, rack, and cleaner, 11.
—— - holder, 18.
Tin, 68
Tongs, 12
Triangle, 13.

Valencies, 101

Washing bottle, 19
Water bath, 20
Wing top, 10.

Zinc, 68.

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